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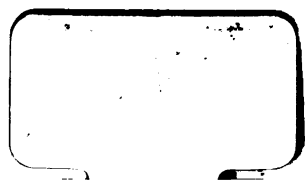
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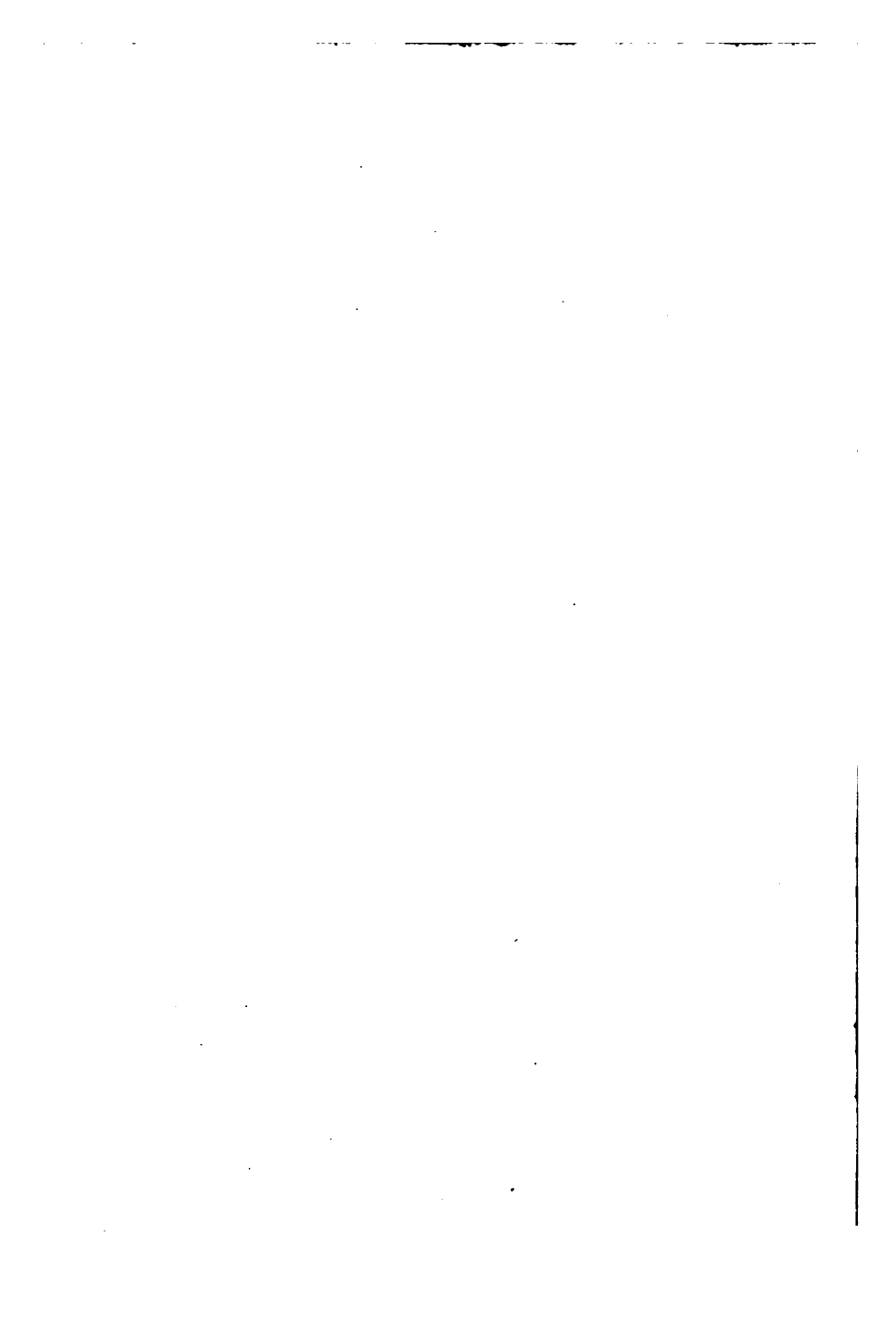
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A CLASS-BOOK
OF
INORGANIC CHEMISTRY:

WITH TABLES OF
CHEMICAL ANALYSIS,

AND DIRECTIONS FOR THEIR USE:

COMPILED SPECIALLY FOR
PUPILS PREPARING FOR THE OXFORD AND CAMBRIDGE MIDDLE-CLASS
EXAMINATIONS, AND THE MATRICULATION EXAMINATION OF
THE UNIVERSITY OF LONDON.

BY

D. MORRIS, B.A.

LATE TEACHER OF CHEMISTRY IN LIVERPOOL COLLEGE.

NEW AND ENLARGED EDITION.



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PREFACE TO NEW EDITION.

IN the present Edition the whole of the work has been carefully revised, and a considerable amount of new matter has been added. The additions include chapters on the Compounds of Carbon and Nitrogen, Light and Spectrum Analysis, the Diffusion of Gases, &c., while the number of Examination Questions on the various parts of the work has been largely increased. The chapter on the Solution of Chemical Problems involving arithmetical operations has been re-written and extended, and a number of selected typical problems are appended to it for solution by the student.

In that portion of the work treating of Practical Chemistry several important changes have been made. The list of apparatus required for simple Qualitative Analysis is new, and will, it is believed, be found to be a tolerably complete list of all that is required for that most essential part of modern chemistry—laboratory practice. Table I., for the preliminary examination of substances in the dry way, has been re-arranged with a view to greater simplicity and accuracy, while Table V., for the detection of the more important acids, is quite new.

The changes made in the present edition are calculated to render the work more useful for the purpose for which it was originally compiled, viz., to furnish students preparing for the University Examinations with a simple yet complete text book.

LIVERPOOL.

September, 1880.

PREFACE TO FIRST EDITION.

THE following pages are simply enlarged notes of the most important facts of Inorganic Chemistry, which were originally culled from the best modern books on this subject for the use of the pupils under my charge. I have long felt the need of an Elementary Text-Book which should satisfy the requirements of the "Middle-Class Examinations;" and, to meet this want, I have compiled, in addition to the main facts of the science, Tables of Analysis of simple compound substances, with directions for their use, and Rules for sundry Arithmetical Exercises.

I strongly recommend to beginners in Qualitative Analysis the use of Table on page 180, on account of the simplicity of the method, as well as for the reason there stated.

The numerous Questions found in the work have been selected from the examination papers of Oxford, Cambridge, and London Universities, and the Science and Art Departments. The questions of the latter are denoted by * : those of the Universities, by the initial letter of each respectively.

My thanks are due to Dr. Carter, of the Liverpool School of Science, for many valuable hints and suggestions.

D. M.

LIVERPOOL,
Aug., 1870.

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INTRODUCTION.

CHEMISTRY has been defined as "the science which treats of the composition of all kinds of matter, and of those changes in composition which result from the action, either of different kinds of matter upon each other, or of external forces upon one and the same kind of matter."—(Frankland).

It is usual, for the sake of convenience, to divide Chemistry into two parts, termed respectively, *Organic* Chemistry, and *Inorganic* Chemistry. This division of Chemistry was at one time believed to be of very great importance, but it is now seen to be purely arbitrary, and the same laws and principles govern both branches. Organic Chemistry deals for the most part with substances formed in connection with plants and animals (organised beings); while Inorganic Chemistry deals with those mineral substances which form part of the crust of the earth.

Most substances met with in nature are what are termed by the chemist *compound substances*, or *compounds*, that is, substances which can, by suitable processes, be decomposed, or separated, into substances which differ from each other and from the original substance. Thus, water is a compound, and can be decomposed into two gases, known as oxygen and hydrogen, both of which possess properties distinct from each other and from the water from which they were derived. But up to the present time no one has succeeded in decomposing oxygen or hydrogen, and such substances are therefore termed by the chemist *elementary substances*, or *elements*. At the present time 64 elements are recognised, and the vast number of different substances met with in nature are formed by the combination of these elements with each other. A comparatively small number only of the elements are of practical import-

ance. Elementary substances are usually divided into *metals* and *non-metals* (sometimes termed *metalloids*), but there is in reality no sharp line of division between the two classes; and some elements—tellurium for example—are placed sometimes in one class, and sometimes in the other. The metals are, with one exception, (mercury or quicksilver), solid at ordinary temperatures, and as a general rule are opaque, and have a characteristic lustre (called the metallic lustre), and a high specific gravity, and are good conductors of heat and electricity.

Matter is believed to consist of very minute portions, which are called *molecules*, and a molecule is defined as the smallest portion of any substance which can exist in a free state. These molecules, or minute portions of matter, are regarded as unalterable by the action of physical forces, and the solid, liquid, or gaseous condition of a body depends entirely upon the state of aggregation of the molecules. Thus, when the molecules are held together in fixed positions by the force of cohesion, we have a solid body; when they are so loosely held together that they are able to move about easily among each other, we have a liquid; whereas when there is no cohesion between the molecules, and they obey the repulsive force which tends to make them expand indefinitely, we have a gas.

The molecules, even of elementary bodies, consist of still smaller portions of matter termed *atoms*, and these atoms are assumed to be indivisible. An atom may be defined as the smallest portion of an element which enters into chemical combination. A molecule may therefore be regarded as a cluster or system of atoms, held together by the force known as *chemical affinity*. The molecules of an elementary substance are each composed of atoms similar in all respects, and, as a general rule, each molecule contains two atoms of the element. Thus each molecule of hydrogen contains two atoms of hydrogen, and each molecule of nitrogen two atoms of nitrogen. A molecule of a compound substance contains two or more different kinds of atoms, but each molecule of the compound

contains the same number of atoms, and of the same kind. Thus a molecule of water contains two atoms of hydrogen united with one atom of oxygen; a molecule of ammonia contains one atom of nitrogen and three atoms of hydrogen; a molecule of sulphuric acid consists of one atom of sulphur, two atoms of hydrogen, and four atoms of oxygen. It is the province of the science of Chemistry to ascertain what is the constitution of the molecules of bodies, and the changes in constitution which the molecules undergo when placed under certain conditions.

It is found that when chemical combination takes place, *the elements unite only in fixed and definite proportions by weight*. To each element there is assigned a number, termed its *atomic weight* or *combining weight*, and this number expresses the relative weight of an atom of the element. It must be clearly understood that it is the *relative*, and not the *absolute*, weight which is thus expressed. And since the weight of the smallest quantity (atom) of hydrogen which enters into combination is less than that of any other element, hydrogen is taken as the standard in this respect, and the weight of an atom of hydrogen is taken as the unit. The atomic or combining weight of oxygen is 16, which expresses the fact that the smallest portion of oxygen met with in chemical combination weighs 16 times as much as the smallest portion of hydrogen, whatever the absolute weight of the latter may be. In a molecule of water, therefore, the one atom of oxygen weighs eight times as much as the two atoms of hydrogen together, and thus water is said to consist of eight parts of oxygen by weight, to one part of hydrogen. Sometimes the *molecular weight* of an element is given instead of its atomic weight, and it is found simply by multiplying the atomic weight by the number of atoms in the molecule. Thus the molecular weight of hydrogen is 2, of oxygen 32, &c. In the same manner the molecular weight, or combining weight, of a compound substance is equal to the sum of the atomic weights of

its constituent atoms; for example, the molecular weight of water is 18, of sulphuric acid 98. This depends on the fact that in chemical combination there is no destruction of matter, or loss of weight, but the weight of the compound is equal to the sum of the weights of its constituents.

The names of the chemical elements have been formed on no definite principle, but it will be observed that a large proportion of the names of the metallic elements end in *ium*. For the sake of convenience in chemical writings, it is agreed to represent an element by a symbol, which is in some cases merely the initial letter of the name (English or Latin) of the element; but in other cases it consists of the first letter combined with a subsequent letter, for the sake of distinction. Thus Hydrogen is represented by the symbol H, Phosphorus by the symbol P, Magnesium by the symbol Mg, and so on. The symbol for the molecule of a compound body is formed by writing the symbols of its constituent atoms, in succession. If two or more atoms of the same element are present in the molecule, the symbol for the element is written only once, and a small figure is added to denote the number of atoms present; when no figure is expressed, one atom is understood. Thus H_2O is the symbol for a molecule of water, and it expresses the fact that such a molecule is composed of two atoms of hydrogen combined with one atom of oxygen; H_2SO_4 is the symbol for sulphuric acid. When it is desired to express more than one molecule of a substance, a large figure is placed before the symbol for the molecule; thus $2\text{H}_2\text{O}$ represents two molecules of water.

The following table gives the names, symbols, and atomic weights of the elements at present known. The words placed in brackets after some of the ordinary names, are the old names of the elements, in nearly all cases from the Latin, and are given to explain the derivation of the symbol. The names of the non-metallic elements are printed in capitals (thus, HYDROGEN); of the important metals, in ordinary type

(thus, Iron); and of the rarer metals, in italics (thus, *Lithium*). The atomic weights of the elements given are in most cases only approximately correct, but they are sufficiently correct for all practical purposes: the atomic weights of some elements, so far as they have been ascertained, extend to several places of decimals.

Names.	Symbols.	Combining Weights.	Names.	Symbols.	Combining Weights.
Aluminium	Al.	27.5	<i>Molybdenum</i>	Mo.	96
Antimony	Sb.	122	Nickel	Ni.	58.6
(Stibium)			<i>Niobium</i>	Nb.	94
Arsenic	As.	75	NITROGEN	N.	14
Barium	Ba.	137	<i>Osmium</i>	Os.	199
Bismuth	Bi.	210	OXYGEN	O.	16
BORON	B.	11	<i>Palladium</i>	Pd.	106
BROMINE	Br.	80	PHOSPHORUS	P.	31
<i>Cadmium</i>	Cd.	112	Platinum	Pt.	197
<i>Cæsium</i>	Cs.	133	Potassium		
Calcium	Ca.	40	(Kalium)	K.	39
CARBON	C.	12	<i>Rhodium</i>	Rh.	104
<i>Cerium</i>	Ce.	141	<i>Rubidium</i>	Rb.	85
CHLORINE	Cl.	35.5	<i>Ruthenium</i>	Ru.	104
Chromium	Cr.	52.5	SELENIUM	Se.	78
Cobalt	Co.	58.6	SILICON	Si.	28
Copper (Cuprum)	Cu.	63.5	Silver		
<i>Didymium</i>	Di.	147	(Argentum)	Ag.	108
<i>Erbium</i>	Er.	169	Sodium		
FLUORINE	F.	19	(Natrium)	Na.	23
<i>Gallium</i>	Ga.	70	Strontium	Sr.	87
<i>Glucinum</i>	Gl. or		SULPHUR	S.	32
(or <i>Beryllium</i>)	Be.	9	<i>Tantalum</i>	Ta.	182
Gold (Aurum)	Au.	196.2	TELLURIUM	Te.	128
HYDROGEN	H.	1	<i>Thallium</i>	Tl.	204
<i>Indium</i>	In.	113	<i>Thorium</i>	Th.	231
IODINE	I.	127	Tin (Stannum)	Sn.	118
<i>Iridium</i>	Ir.	197	<i>Titanium</i>	Ti.	48
Iron (Ferrum)	Fe.	56	<i>Tungsten</i>		
<i>Lanthanum</i>	La.	139	(Wolfram)	W.	184
Lead (Plumbum)	Pb.	207	<i>Uranium</i>	U.	240
<i>Lithium</i>	Li.	7	<i>Vanadium</i>	V.	51
Magnesium	Mg.	24	<i>Yttrium</i>	Y.	92
Manganese	Mn.	55	Zinc	Zn.	65
Mercury			<i>Zirconium</i>	Zr.	90
(Hydrargyrum)	Hg.	200			

Tellurium is sometimes classed as a metal, and Arsenic is frequently included among the non-metallic elements. Quite recently the discovery of several new elements (metals) has been announced, but they are not sufficiently well established to be included in the above list.

The word *density* is of frequent occurrence in chemistry, and means the relative quantity of a given substance contained in a certain space—a cubic inch for example—compared with the quantity of matter contained in a cubic inch of water, which is taken as the standard for liquids and solids; or of hydrogen or air, one of which is taken as the standard of reference for gases. Since the quantity of matter in any body is determined by the weight of the body, density is practically the same thing as specific gravity. By the *specific gravity* of a substance is meant the relative weight of a certain volume of the substance, compared with the weight of an equal volume of some other substance chosen as a standard. Pure water is chosen as the standard with which to compare liquids and solids, and when it is said that sulphuric acid has a specific gravity of 1·8, and gold a specific gravity of 19·3, it is meant that if equal volumes of water, sulphuric acid, and gold were taken and weighed, the respective weights would be in the proportion of the numbers 1 : 1·8 : 19·3. Air is usually taken as the standard with which any gas is compared as regards its specific gravity.

In works on Chemistry, the metric or French system of weights and measures is very frequently employed, on account of its simplicity. A table, comparing the English and French systems, will be found in the Appendix.

THE NON-METALLIC ELEMENTS.

OXYGEN.

Symbol O. Combining weight 16.

This gas was discovered by Dr. Priestley, in England, in 1774, and Scheele, in Sweden, in the year 1775. It was called "Oxygen" (**Acid producer**) because at that time all acids were supposed to contain this element. It is the most widely diffused substance in nature: it forms one-fifth of the air by volume, and eight-ninths of water by weight, and it has been estimated to form nearly one-half of the solid portion of the earth. It is found in all plants and animals, in rocks and earths. When free, it is always observed in the state of gas.

Preparation. Some substance which contains it, and readily parts with it, is placed in a flask or tube, and heat applied. The substance usually taken is **potassium chlorate** (chlorate of potash) KClO_3 . On the application of heat, this soon melts and gives off bubbles of oxygen gas, which may be collected by the displacement of water in inverted glass jars. What remains in the flask is **potassium chloride** (chloride of potassium) KCl . The chlorate yields 39.2 per cent. of its weight of oxygen; and 1 oz. gives nearly 2 gallons of gas.

The change may be shown in symbols thus:—

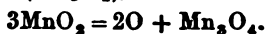


The sign + in all chemical equations means "together with,"

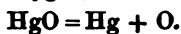
The sign = may be read "gives."

Oxygen comes off more quickly when a little powdered **Manganese dioxide** (black oxide of **Manganese**), MnO_2 , is added to the potassium chlorate; the manganese dioxide, however, undergoes no change itself, and its action is not thoroughly understood.

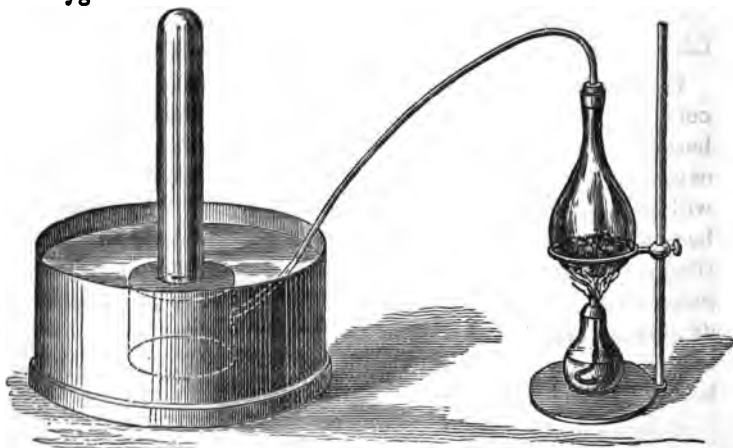
This gas is also prepared in large quantities by heating the **black oxide of Manganese** (MnO_2), which yields 12.3 per cent. of oxygen by weight. The black oxide is converted into a reddish-brown oxide (Mn_2O_3), thus:—



When **Mercury** is heated for a long time in the air, it combines with the oxygen of the air and forms a red substance, which is known as the **red oxide of Mercury**, and is expressed by the symbol HgO . This red oxide of **Mercury**, heated in a glass tube, gives off its oxygen, thus:—



It was from this compound that Dr. Priestley first obtained oxygen.



Properties. Oxygen has a density slightly greater than atmospheric air, its density being 1.1056, that of air being

1.000. Oxygen is 16 times heavier than hydrogen, volume for volume. It has no colour, smell, or taste. It is slightly soluble in water, 100 cub. inches of water at 0° C. dissolve 4 cub. inches of the gas. Oxygen is always met with in the state of a gas, and was for a long time classed along with hydrogen and nitrogen as a "permanent" gas. In December 1877, MM. Pictet and Cailletet independently succeeded in reducing these so-called permanent gases to a liquid condition, by the combined action of very great cold and pressure.

Oxygen is essential to the existence of animal life: it forms what used to be called the **vital air** of the atmosphere. It has the power of combining with every element except one (Fluorine). One of its most striking properties is its power of supporting combustion. A red hot splinter of wood plunged into a jar of the gas bursts into flame. Phosphorus burns in it with great brilliancy; sulphur and charcoal do the same. Many metals burn vividly in it; for example, a piece of thin iron wire, formed into a spiral, and tipped with sulphur, will, when lighted and plunged into the gas, burn brilliantly. Zinc foil may be burned in the same way.

Combustion in air and in oxygen is exactly the same thing, only that, as oxygen forms but one-fifth part of the air, combustion in the air is neither so rapid nor so brilliant as in pure oxygen. During the process of combustion, the oxygen unites with the burning substance and forms an **oxide**. Thus the burning wire forms with oxygen the **black oxide of iron** (Fe_3O_4): Phosphorus forms **phosphorus pentoxide** (P_2O_5): Sulphur forms **sulphur dioxide** (SO_2): Charcoal forms **carbon dioxide** or **carbonic acid** (CO_2). The act of combination of any substance with oxygen is called **Oxidation**, during which heat is always, and light is frequently, given off.

These oxides are compounds, and, as is the case with all chemical compounds, their weight is exactly equal to the sum of the weights of the elements composing them.

OXIDES are very numerous and important. They are conveniently divided into three classes. 1. **Acid oxides**, or acid-forming oxides. 2. **Basic oxides**. 3. **Neutral oxides**.

Acid Oxides are those which, when soluble in water, have a sour taste, turn vegetable blue colours (litmus, for example) to a bright red, and browns (turmeric) to yellow. Sulphur trioxide (SO_3), Nitrogen pentoxide (N_2O_5), and Silica or Silicic acid (SiO_2), are examples of this class. The last named oxide (silica) is not soluble in water, and therefore has no sour taste, and does not affect the colouring substance, but its acid property is shown by the fact that it unites with certain basic oxides.

Basic Oxides have properties the very opposite of Acid Oxides. They change vegetable reds to blue or green, and yellows to brown; they neutralize the strongest acids, and combine with acids to form compounds called **salts**. Potash (Oxide of Potassium, K_2O), and Soda (Oxide of Sodium, Na_2O) are examples of this class.

Neutral Oxides are those which have no sour taste, no action upon colouring matter, and do not combine with acids or bases. Water (Oxide of Hydrogen), and Carbon monoxide (CO), are examples of this class.

It frequently happens that a body combines with oxygen in different proportions, and forms a series of oxides. There are, for instance, seven different oxides of the metal Manganese. In order to distinguish them, Latin (or Greek) numerals are prefixed to the term oxide. Thus, if the oxide is formed with one atom of oxygen, it is called a **monoxide** or **protoxide**; if with two atoms of oxygen, it is called a **dioxide** or **binoxide**; if with three of oxygen, a **trioxide** or **teroxide**, and so on. The highest oxide is often called a **peroxide**. Sometimes two atoms of a metal unite with three of oxygen, then the name **sesquioxide** is frequently given,—*sesqui* means one and a half. Any compound containing less oxygen than a **protoxide** is called a **suboxide** (*e.g.* Ag_4O , Silver suboxide; Ag_2O , Silver monoxide or protoxide).

Acids are distinguished in the following way. The name of the acid which has the largest proportion of oxygen, ends in the syllable **ic**. An acid containing less oxygen has a name ending in **ous**. Thus in the two compounds, H_2SO_4 , and H_2SO_3 , the first contains one atom of oxygen more than the second; hence the former is called Sulphuric Acid, and the latter Sulphurous Acid.

When Acids and Basic oxides combine with each other, they form what are called **Salts**. When acids whose names end in **ic** form salts, this termination is changed into **ate** to express the name of the salt. When sulphuric acid combines with iron, the salt thus formed is called sulphate of iron, or ferrous sulphate (FeSO_4).

If the name of the acid ends in **ous**, the name of the salt ends in **ite**. Sulphurous acid combines with baryta and forms the salt sulphite of baryta, or barium sulphite (BaSO_3).

The names of some salts end in the syllable **ide**. This termination is used when **two elementary bodies** unite to form a salt. Sulphur and sodium unite to form the salt sulphide of sodium (Na_2S). Similarly we have ox-ide, hydr-ide, chlor-ide, fluor-ide, &c.

Care must be taken to distinguish the **ites** from the **ates**, and the **ides** from the **ites**.

OZONE.—When electric sparks are sent through the air or dry oxygen, a faint smell is perceived. To the body which produces this smell the name **ozone** is given, from a Greek word, meaning “to give off an odour.” Ozone is supposed to be oxygen in a more active state after undergoing condensation. Experiment shows that 3 volumes of ordinary oxygen condense to form 2 volumes of ozone, and it is believed that a molecule of ozone contains 3 atoms of oxygen.

Ozone is formed in small quantities whenever water is decomposed between platinum plates by electricity. It may also be formed chemically, by allowing a stick of phosphorus to

hang in a bottle filled with moist air ; and by the action of strong sulphuric acid upon the permanganate of potassium. It possesses powerful bleaching properties, and is present in small quantities in the atmosphere, but scarcely ever in large towns. It is insoluble in water, and in solutions of the acids and alkalies, but is absorbed by a solution of potassium iodide (iodide of potassium). The most delicate test of ozone is potassium iodide (KI), from which it immediately sets iodine free. If a piece of paper be dipped in a solution of potassium iodide and starch paste, and then exposed to the action of ozone, the paper becomes blue, owing first to the liberation of the iodine, and, secondly, to the formation of a blue compound of iodine and starch.

MEASUREMENT OF TEMPERATURE OF BODIES.

THERMOMETERS.—Changes in the temperature of bodies are measured by instruments called **thermometers**. These are closely sealed glass tubes containing a liquid, usually Mercury, which expands or contracts in volume according as its temperature is increased or diminished. Upon the tube are marked the points at which the mercury stands, when the tube is placed in freezing or in boiling water respectively. These are termed the “fixed points,” and the distance between these points is graduated according to one of the following scales :—
1. Fahrenheit’s scale ; 2. Centigrade scale ; 3. Reaumur’s scale. Fahrenheit’s scale is the one most commonly used in England. According to this, the freezing point of water is marked at 32 degrees, and the boiling point at 212 degrees, the distance between the points being divided into 180 equal parts, termed degrees. In the Centigrade thermometer these points are marked at zero (0 degrees) and 100 degrees respectively ; while, according to Reaumur’s scale, the same points are marked at zero and 80 degrees respectively. The Centigrade thermometer is used in France, and is now generally employed by scientific men in England. The temperatures mentioned in the following

pages are usually expressed on the Centigrade scale. It is very easy to convert temperatures expressed according to one scale into corresponding temperatures on the other scales. Thus :— Since the distance between freezing point and boiling point is on the Centigrade scale divided into 100 degrees, and on the Fahrenheit scale 180 degrees, it follows that 100 Centigrade degrees are equal to 180 Fahrenheit degrees, or $5^{\circ}\text{C.} = 9^{\circ}\text{F.}$ Suppose we want to know what degree of Fahrenheit's scale corresponds to 60°C. Since $5^{\circ}\text{C.} = 9^{\circ}\text{F.}$, therefore $60^{\circ}\text{C.} = 108^{\circ}\text{F.}$; but as Fahrenheit's scale commences with 32° at the freezing point, that number must be added to 108° . Hence the answer will be 140°F. On the same principle, the Fahrenheit may be expressed in terms of the Centigrade. To express 60°F. on the Centigrade scale, subtract 32° , then, since $9^{\circ}\text{F.} = 5^{\circ}\text{C.}$; $28^{\circ}\text{F.} = 28 \times \frac{5}{9} = 15\frac{5}{9}^{\circ}\text{C.}$

The rule for conversion may be expressed thus :—

From C. to F. Multiply by 9, divide by 5, and add 32.

„ F. to C. Subtract 32, multiply by 5, and divide by 9.

EXPANSION OF GASES BY HEAT.

It has been found, by careful experiments, that gases expand $\frac{1}{273}$ part of their volume at 0°C. , for every increase of 1°C. Thus :—A certain quantity of gas, which at 0°C. occupies 273 volumes, will at 1°C. occupy 274 (*i.e.* $273 + \frac{1}{273}$ of 273), and at 10°C. it will occupy 283 (*i.e.* $273 + \frac{10}{273}$ of 273), &c.

The fraction $\frac{1}{273}$ expressed as a decimal, is .003665. If 100 measures of a gas be measured at 0°C. , and then heated to 1°C. , they become 100.3665. Similarly, 100.3665 volumes of gas at 1°C. , become 100 volumes at 0°C. Remembering the fact stated above, if the volume of any gas at 0°C. be given, its volume at any degree of heat and the same atmospheric pressure may be readily found.

RULE FOR CORRECTION OF GASES FOR TEMPERATURE.

As 273 + given temperature (Centigrade) is to 273 + required temperature, so is the given volume to the required volume.

RELATION OF VOLUME OF GASES TO PRESSURE.

The volume of a gas increases or decreases according to the pressure to which it is subjected. For example, a certain weight of gas occupies 100 cubic inches at the ordinary pressure of the air; if the pressure be doubled, the volume of the gas will be 50 cubic inches, or one-half of its former volume; if the pressure be diminished one-half, the volume will be 200 cubic inches, or double. The relation of the volume of a gas to the pressure is stated in the following law, called **Boyle's**, or **Marriott's Law**,—"The volume of a gas is inversely as the pressure."

From the above statement it is evident that, if the volumes of gases are to be compared, some standard measure of temperature and pressure must be taken. For this purpose it is agreed to take the temperature at 32° F. or 0° C., and the standard pressure of 30 inches, or 760 millimetres (m.m.) of Mercury.

The following exercise, worked out in full, will serve to illustrate, not only the application of the rules concerning the volume of a gas just given, but also the method of treating those problems in which it is required to ascertain what weight of a chemical substance may be obtained from a compound containing it. Suppose it is required to find what weight of Potassium Chlorate will yield 20 litres of oxygen, the gas being measured at a temperature of 15° C., and under a pressure of 750 m.m. of Mercury.

$$\text{Potassium Chlorate (KClO}_3\text{)} \left\{ \begin{array}{ll} \text{K (potassium)} & 39 \text{ parts by weight.} \\ \text{Cl (chlorine)} & 35.5 \text{ } \\ \text{3O (oxygen)} & 48 \text{ } \end{array} \right. \quad \begin{array}{l} \\ \\ \text{,,} \end{array}$$

122.5

Therefore, 122.5 parts by weight of Potassium Chlorate will yield 48 parts by weight of oxygen.

Again, 20 litres of oxygen at 15° C. and 750 m.m. will

become, under standard conditions of temperature and pressure, (0° C. and 760 m.m.) by rules just given :—

$$20 \times \frac{750}{760} \times \frac{273}{288} = 18.7 \text{ litres.}$$

One litre of oxygen at 0° C. and 760 m.m. weighs 1.4336 grammes ; therefore 18 litres will weigh 26.808 grms.

To furnish this quantity of oxygen, the quantity of Potassium Chlorate required is :—

$$\frac{26.808 \times 122.5}{48} = 68.4 \text{ grms.}$$

QUESTIONS ON OXYGEN, &c.

1. What class of bodies do we call oxides ? Name some oxides of an acid nature, and some basic oxides. (L.)
2. Describe and explain the preparation of Oxygen from Potassium Chlorate (Chlorate of Potassium); and name the substances which are formed when Carbon, Hydrogen, Sulphur, and Phosphorus are respectively burned in an excess of Oxygen. (L.)
3. What weight of Oxygen may be obtained by the decomposition of 53 grains of Chlorate of Potassium ? (O.)
4. What is Ozone ? Under what circumstances is it formed ? What are its properties ? (O.)
5. How many grammes of Oxygen can be separated from 100 grammes of Bin oxide of Manganese by heat ? And how many by the action of Sulphuric Acid ? *
6. Write out Boyle's Law. 20 litres of a gas are measured at 0° C. and 720 m.m. pressure. Required the volume at 760 m.m., the temperature being the same. *
7. 40 cubic centimetres of a gas are measured at 20° C. How many will it measure at 30° C. ? *
8. 40 cubic centimetres of air are measured at 15° C. and 750 m.m. pressure. Required the volume of air at 25° C. and 800 m.m. pressure.
9. A gas measures 117.9 cubic inches at $9^{\circ}.9$ C. and 709 m.m. pressure. Required its volume at 0° C. and 1 metre pressure. *
10. 600 cubic centimetres of gas at 10° C. is warmed until it becomes 730 cubic centimetres. Through what number of degrees C. has the gas been heated ? *

11. A grain of alcohol is converted into vapour of the temperature of 100°C ., under a pressure of 741 m.m. Mercury; what volume does the vapour occupy? (O.)

12. How is a thermometer graduated? Describe the three scales in common use. Express -30°Fah. and $+327^{\circ}\text{Fah.}$ in terms of the Centigrade scale. (L.)

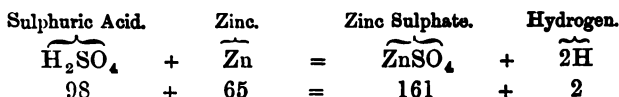
HYDROGEN.

Symbol H. Combining weight 1. Sp. gr. (compared with air) .0692.

Hydrogen was discovered by Cavendish in 1766. It is rarely found in nature uncombined. It is one of the constituents of water, of which it forms one-ninth part by weight. It is found as a constituent of all plants and animals. It receives its name from two Greek words, meaning "water producer."

Preparation. Hydrogen may be obtained from the decomposition of water by means of the metals potassium and sodium. For this purpose a small fragment of the metal is dropped into a vessel of water, the metal floats about, decomposes the water, uniting with the oxygen to form potash (K_2O) or soda (Na_2O), and liberating the hydrogen. The potash or soda formed at once dissolves in the water. The hydrogen may be caught in a jar, which is filled with water and inverted with its mouth dipping into the water in the vessel over the place where the hydrogen is being liberated. Another method of preparing hydrogen is to pass vapour of water (steam) through a tube in which fragments of iron are heated to redness. The red-hot iron decomposes the water, uniting with the oxygen to form black oxide of iron (Fe_3O_4), and the hydrogen is set free.

Hydrogen is usually made by pouring dilute sulphuric (or hydrochloric) acid on to zinc clippings; the acid dissolves the zinc to form sulphate (or chloride) of zinc, setting hydrogen free, thus:—



The numbers are the chemical equivalents or combining weights of the substances comprised in the equation, and they show for example that for every 65 parts by weight of zinc, 2 parts of hydrogen are obtained.

One oz., or 15·44 grammes, will disengage 629 cub. in., or about 10 litres of hydrogen from water. Water may be decomposed into its constituent elements by sending a strong current of electricity through water (slightly acidulated by sulphuric acid), between two platinum plates. Oxygen gas is given off in bubbles at one plate, and hydrogen at the other plate.

Properties. Hydrogen is the lightest body in nature, the density of oxygen being 16 times that of hydrogen, and the density of the air about $14\frac{1}{2}$ times as great. 100 cub. in. of it at 60° F. and 30 in. Bar. weigh 2·14 grains; or, according to the metrical system, 1 litre at 0° C. and 760 m.m. Bar. weighs 0·0896 grammes.

Owing to its lightness, it has often been used to fill balloons. It is a bad conductor of sound; a bell rung in a jar of the gas gives a very feeble sound. It is slightly soluble in water, 100 volumes of water absorb about 2 volumes of the gas. Hydrogen has recently been liquefied by the combined action of extreme pressure and cold.

Hydrogen burns in the air with a pale flame, giving out much heat, but little light. During the combustion, it unites with oxygen, and forms vapour of water. If a cold vessel be held over a burning jet of the gas, it soon becomes covered with drops of water.

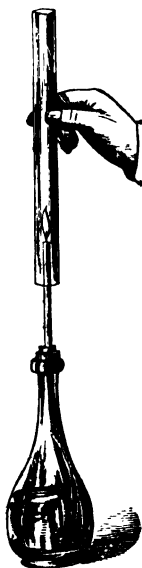
If two volumes of hydrogen be mixed with one volume of oxygen, and a light applied, the gases immediately combine with a loud explosion, and form water. Strong jars are necessary for this experiment, and it is well to wrap a cloth round the vessel containing the mixture. On account of the explosive property of hydrogen when mixed with oxygen, care is required in collecting the gas. Time should be allowed for the expulsion of the air from the vessels in use.

If a glass tube, one or two feet in length, be held over a burning jet of the gas, a musical note is produced. This sound is caused by the rapid explosions of hydrogen and oxygen in the confined space of the tube.

Though the flame of this gas has a feeble light, yet, when mixed with oxygen and made to play upon a piece of lime, it produces a most brilliant light, which on great elevations can be seen nearly 100 miles off on a clear night.

Hydrogen forms numerous compounds with other bodies, but it does not equal oxygen in this respect.

It is readily absorbed by many porous bodies, particularly by the metals platinum and palladium. If a piece



of spongy platinum be held in a jet of the gas as it escapes into the air, the cold platinum will become red hot. Some chemists regard hydrogen as a metal, and there is very much to be said in favour of this view; but as the evidence so far is incomplete, it is for the present retained among the non-metallic elements.

COMPOUNDS OF OXYGEN AND HYDROGEN.

There are two compounds of Oxygen and Hydrogen, viz:—

1. *Water*. Symbol H_2O . Combining weight 18.
2. *Peroxide of Hydrogen*. Symbol H_2O_2 . Combining weight 34.

WATER.—Cavendish discovered, in 1781, that water consisted of the two gases, hydrogen and oxygen, in the proportion of two volumes of hydrogen to one of oxygen: the proportion by weight in which oxygen and hydrogen are respectively combined to form water, is as 16 : 2. A claim to the same discovery has been made on behalf of James Watt. Water may be resolved into its two elements by sending an electric current through it, when acidulated with a little sulphuric acid.

This experiment is performed as follows:—Two test tubes, filled with water, are inverted over two small platinum plates, in a glass vessel containing acidulated water. The plates are then connected with a voltaic battery, and as the current passes through the liquid between the plates, the water is decomposed, and the disengaged gases ascend into the tubes respectively. If the tubes are graduated, it will be seen that the volume of hydrogen disengaged will be double that of the oxygen: the proportions are not exactly shewn, in consequence of the greater solubility of the latter gas.

The **Eudiometer** is an instrument which may be used to shew the proportions in which the two gases combine to form water. It is simply a strong glass tube, accurately graduated, closed at one end, through which two platinum wires are fixed. The tube is first filled with mercury, and inserted in a vessel containing the same liquid. Hydrogen gas, say to the extent

of 100 volumes, is passed into the tube, and then some oxygen, say 100 volumes. An electric spark is sent through the mixture, between the platinum wires, vapour of water is formed, and the mercury ascends, until impeded by the gas remaining in the tube. This is pure oxygen, which the eudiometer will shew to be 50 volumes. The other 50 volumes of oxygen have united with the 100 volumes of hydrogen to form the water which is condensed on the sides of the instrument. This experiment therefore proves the truth of the previous statement, that the composition of water by volume is two parts of hydrogen to one of oxygen (H_2O).

Pure water is a colourless and transparent liquid, which has neither taste nor smell. It freezes or solidifies at $0^\circ C.$ ($32^\circ F.$) forming the solid called ice, which has a somewhat less density than water, on account of the expansion which occurs on solidification. At $100^\circ C.$ ($212^\circ F.$), water boils, forming a vapour called steam, which, under the ordinary atmospheric pressure, occupies a volume about 1700 times that of the water from which it was derived. The temperature of the boiling point of water varies with the pressure to which it is exposed, the temperature just given being that of boiling water under the ordinary atmospheric pressure at the sea level (about 15 lbs. to the square inch). At greater heights in the atmosphere the pressure becomes less, and the boiling point of water is proportionally lower. This fact is sometimes made use of to estimate the heights of places above the sea level.

Water combines with many compounds to form what are called **hydrates**, thus :—potash (K_2O) unites with water to form potassium hydrate ($K_2O + H_2O = 2KHO$); and quick lime, which is calcium oxide (CaO), combines with water to form calcium hydrate (CaH_2O_2). Certain oxides combined with water form the compounds known as acids. These oxides are called **anhydrides**; Sulphur dioxide, or sulphurous anhydride (SO_2), unites with water to form sulphurous acid (H_2SO_3). Compounds devoid of water are termed **anhydrous** (without water).

Water occurs in many crystals, and its presence is essential to the existence of the crystalline form ; in this condition it is known as "water of crystallisation." The crystals of carbonate of soda (washing soda) contain much water of crystallisation, and when these are exposed to the air, the water passes away, leaving the substance in the form of a white amorphous powder.

Water has the property of dissolving a large number of solid substances, and, as has been stated, it can also dissolve oxygen, hydrogen, and other gases. In general its solvent power on solids increases with the temperature, but the reverse is the case as regards the solution of gases. From its great solvent powers, water is never met with absolutely pure. **Distilled water** is the purest form of water used, and is obtained by evaporating water and condensing the steam. Any substances which may have been dissolved in the water do not pass off along with the steam. **Rain water** is water distilled from the sea and other natural collections of water ; in its descent it dissolves gases and other substances from the air to a small extent. **River water** contains more or less matter in suspension, and also in solution, the dissolved matters varying in nature and amount with the nature of the rocks over which the water has passed. Carbonate of lime is a substance found very commonly in river water ; it gives to water the quality of "hardness." **Spring water** usually contains a considerable amount of solid matter in solution :—*e.g.*, carbonate of lime, sulphate of magnesia, &c. Some varieties of spring water, containing compounds of sulphur and iron, and other substances, are used medicinally. **Sea water** usually contains about $3\frac{1}{2}$ per cent. by weight of substances in solution, the one which is present in by far the largest quantity being common salt (chloride of sodium). An analysis of the water of the English Channel gave, in 1000 parts:—Water, 964·7; Sodium chloride, 27·0; Magnesium chloride, 3·6; Magnesium sulphate, 2·3; Calcium sulphate, 1·4; with smaller quantities of Calcium carbonate and other substances.

HYDRIC PEROXIDE, or **Peroxide of Hydrogen** (H_2O_2), is not found in nature. It differs from water in possessing twice as much oxygen; and because it easily gives off half its oxygen, and turns to water, it is called sometimes **oxygenated water**.

It is prepared by acting on barium peroxide (BaO_2) with hydrochloric acid; barium chloride (BaCl_2), and hydric peroxide (H_2O_2), are formed.

This latter compound possesses bleaching properties.

COMBINING POWERS OF THE ELEMENTS.

An examination of a small number of chemical compounds reveals the fact that the different elements have not all the same value in a chemical combination; and that when one element replaces another in a compound, the number of atoms of the two elements are frequently different. Thus it requires two atoms of hydrogen to take the place of one atom of oxygen. Again, when two elements combine together, it appears that in many cases one atom of an element (hydrogen, for example) is not able to satisfy the chemical affinity of one atom of another element. We see this in the case of water, for example; an atom of oxygen has its chemical affinity satisfied by nothing less than two atoms of hydrogen, whereas an atom of chlorine is satisfied by combination with one atom of hydrogen, while an atom of carbon requires four atoms of hydrogen. This property of an element is sometimes spoken of as its **combining power**, **quantivalence**, or **atomicity**. And the number of atoms of hydrogen or chlorine (which is equivalent to hydrogen in this respect) with which an element is capable of combining, is the measure of its combining power or atomicity.

(1.) Those elements which combine with one equivalent of hydrogen, and which, when they substitute each other in chemical compounds, do so atom for atom, are called **monovalent**, or **monads**. The monad non-metallic elements are **Hydrogen, Chlorine, Bromine, Iodine, Fluorine**.

(2.) Those elements whose atoms are equivalent in combin-

ing power to two atoms of hydrogen, or to two atoms of any monad element, are termed **dyads**, and among non-metallic elements we have **Oxygen, Sulphur, Selenium, Tellurium**.

(3.) **Boron, Nitrogen, and Phosphorus** are termed **triads**, though by some chemists the latter two are considered as **pentads**. In their hydrogen compounds (NH_3 ammonia— PH_3 phosphoretted hydrogen) they are obviously triads, or trivalent.

(4.) **Carbon and silicon** can each satisfy four atoms of hydrogen, as in CH_4 , marsh gas; SiH_4 siliciuretted hydrogen; and are hence termed **tetravalent**, or **tetrads**.

Sulphur sometimes occurs in combination in such a manner as to justify being called a **hexad**, as in sulphur trioxide, SO_3 .

MEASUREMENT AND WEIGHT OF GASEOUS VOLUMES.

Though the various elements differ in their combining proportions, yet an equivalent quantity of each element, in the proportion of its atomic weight (with the exception of arsenic, phosphorus, and mercury), fills exactly the same space when in the state of gas or vapour, at the same temperature and pressure. Thus, one part by weight of hydrogen gas occupies the same space as 16 parts by weight of oxygen, 14 parts of nitrogen, or 35.5 parts of chlorine. The law is usually stated thus:—"Equal volumes of gases and vapours, under the same conditions of temperature and pressure, contain equal numbers of molecules."

From this fact, if we take a fixed volume of hydrogen as a standard, we shall be able to calculate the absolute weight of a given volume of any gas at the standard temperature and pressure; since **the density of a gas, compared with hydrogen, is exactly the same as its combining weight**. The standard volume-weight unit is 1 litre of pure hydrogen, at 0°C . and 760 m.m. pressure; and the actual weight of this quantity is .0896 grammes. The term **crith** has been given to this standard unit. From this, the weight of a litre of any gas or vapour at the same temperature and pressure, is known by multiplying the **crith** by the combining weight of the element.

1 litre of oxygen weighs $16 \times 0.0896 = 1.4336$ grammes.

1 " nitrogen " $14 \times 0.0896 = 1.2544$ "

In the case of almost all compound gases, the density, compared with hydrogen, is one-half the combining weight. Thus, steam, or water-gas (H_2O), is a compound containing two volumes of hydrogen united with one volume of oxygen; the combining weight is therefore $2 + 16 = 18$. On combination these gases contract to such an extent as to form only two volumes of steam; the density of steam, therefore, compared with hydrogen, will be $\frac{18}{2} = 9$; or, one litre of steam weighs 9 criths = 0.8064 grms.

Since 0.0896 grms. of hydrogen, $0^\circ C$. and 760 m.m. occupy 1 litre.

\therefore 1 grm. " " occupies 11.2 litres.

So 16 grms. of oxygen, or 14 grms. of nitrogen, will occupy this volume. And in general, if we take a number of grammes of an element (except arsenic, phosphorus, and mercury) equal to the atomic weight of the element, the volume occupied by that quantity of the element in the state of gas or vapour at $0^\circ C$. and 760 m.m. will be 11.2 litres. In the case of the majority of compounds, the number of grammes taken to fulfil this condition would be equal to half the combining (or molecular) weight.

From this we are enabled to find the volume of a substance produced in certain decompositions, or conversely. For example, if it be required to find the number of litres produced when a kilogramme of zinc is acted on by excess of hydrochloric acid (HCl), express the decomposition in the form of an equation, thus:—



From the above equation we find that 65 grms. of zinc give off 2 grms. of hydrogen.

Since 65 grammes of zinc give off 22.4 litres of hydrogen.

\therefore 1 gramme " $\frac{22.4}{65}$ "

And 1000 grms. of zinc give $\frac{22.4}{65} \times 1000 = 344.6$ litres.

QUESTIONS ON HYDROGEN.

1. How much zinc must be weighed out in order to obtain hydrogen enough to fill a gasholder of 10 litres capacity? (O.)

2. Describe how you would prepare hydrogen from water without the assistance of an acid; and explain your process. How can a mixture of two volumes of hydrogen and one volume of oxygen be converted into water? (L.)

3. State the composition of water by weight, and by volume; and explain how the latter may be shown. If the elements of water were mixed in a close vessel in the proportion of 1 : 3, and exploded, what would be found in the vessel after explosion? (C.)

4. How much water would be produced from 28 lbs. of oxygen and 5 lbs. of hydrogen; and would either of the elements be in excess? *

5. In what respects does rain-water differ from sea-water? Enumerate the most common constituents of spring-water. How is pure-water obtained from spring-water? (L.)

6. Water is said to boil at the top of Mont Blanc at 86° C., on Mount St. Bernard at 92° C., and in London at about 100° C.? Explain fully the causes of these differences. (L.)

7. In what different proportions by volume does hydrogen ordinarily combine with other elements? Give an example of each proportion. How can you prove the proportion in the case of hydrochloric acid? (C.)

8. How would you prove, by means of experiments, that water is a compound of hydrogen and oxygen? (L.)

9. How may pure water be obtained? Describe what happens when a piece of sodium is thrown into water. How may oxygen gas be liberated from water?

NITROGEN.

Symbol N. Combining weight 14. Sp. gr. 0.972.

This gas is found in a free state in the atmosphere, of which it forms four-fifths by volume. It is found in combination in the substance of both plants and animals, and in many minerals, e.g., saltpetre (nitrate of potash, KNO_3). It derives its name, Nitrogen (nitre producer), because it forms the chief part of nitre, or saltpetre. It is also called *azote* (without life), be-

cause, though not in itself poisonous, animals cannot live in it in the absence of oxygen. It was discovered by Rutherford in 1772.

Preparation. The simplest way of obtaining this gas is to deprive air of its oxygen. For this purpose, a fragment



of dry phosphorus is burned under a glass vessel inverted in water. The phosphorus may be placed in a floating dish. All the oxygen in the vessel will be removed during combustion, and nitrogen will remain. The phosphorus unites with the oxygen to form phosphoric anhydride (P_2O_5), which dissolves in the water. A piece of phosphorus left in an inverted jar of air for 24 hours will absorb all the oxygen. Iron clippings moistened will also remove the oxygen.

Nitrogen may also be prepared by sending a current of chlorine gas through ammonia, but this experiment is attended with some risk; also by sending atmospheric air through a tube containing red-hot copper filings. In the latter case, the oxygen of the air unites with the copper to form black oxide of copper (CuO), and the nitrogen comes off free.

Properties. Nitrogen has no colour, taste, or smell; it is always in the gaseous state when free. It is a little lighter than air, its sp. gravity being 0.97 compared with air. Water dissolves about 2 cub. in. of it in every 100 cub. in. at $0^\circ C$.

It neither burns itself nor supports combustion; it cannot support animal life. It will thus be seen that the properties of nitrogen are for the most part negative, and in the free state it is notable chiefly for its chemical inactivity. United with hydrogen, however, it forms a powerful alkaline compound, **ammonia** (NH_3), whilst when united with oxygen and hydrogen, it forms one of the most important and powerful acids, **nitric acid** (HNO_3).

THE ATMOSPHERE

The atmosphere is a **mechanical mixture** of several gases, of which oxygen and nitrogen are the most abundant. It is called a **mechanical mixture**, because the gases in it are not combined to form a **chemical compound**, but are simply mixed with each other. One proof of this consists in the fact that if we take quantities of the gases composing the air in the proportion in which they there occur, and simply mix them together, we have a mixture possessing all the properties of air. Another proof of this is found in the unequal solubility of the two principal gases (oxygen and nitrogen) in water. Were these two gases combined in the air, it would follow that the air dissolved in water would have precisely the same composition as ordinary air; whereas if the air were simply a mechanical mixture we should expect to find a somewhat larger proportion of the more soluble gas. The latter is found to be the case; oxygen is slightly more soluble than nitrogen, and the air dissolved in water contains a larger proportion of oxygen than does ordinary air.

The two chief gases in the air are oxygen and nitrogen; and, from the examination of air in various parts of the world, it appears that the proportion in which these two gases are present is very constant, 100 parts of air contain:—

	By Weight.	By Volume.
Nitrogen - - - - -	76·9	79·1
Oxygen - - - - -	23·1	20·9
	<hr/> 100· <hr/>	<hr/> 100· <hr/>

The relative quantities of these two gases may be shown by the following experiment. Confine air in a glass tube, marked off into 100 equal parts, by inverting it in water: place a piece of phosphorus in the tube,—the oxygen slowly combines with the phosphorus, and the water ascends 21 parts to take its place, leaving the remaining 79 parts filled with nitrogen.

In addition to **nitrogen** and **oxygen**, the following gases and vapours are also present in the atmosphere :—

Carbonic Anhydride, (Carbonic Acid)—from 3 to 6 parts in 10,000.

Aqueous Vapour—amount dependent upon the climate.

Traces of **Nitric Acid**.

„ **Ammonia**.

„ **Carburetted Hydrogen**.

„ **Sulphuretted Hydrogen, and Sulphurous Acid**.

The presence of oxygen in air is proved by burning bodies being changed into oxides, *e.g.*, by the formation of oxide of copper, when a current of atmospheric air is passed over red-hot copper filings. Its presence in the free state is essential to all life.

The presence of nitrogen may be shown in the manner described in the preparation of that gas. It serves to dilute the oxygen of the air, and thus make it suitable for living beings.

The presence of carbonic acid is shown by exposing a vessel containing lime water to the air. After a while a deposit of chalk (carbonate of lime) will be found in the vessel, in consequence of the union of the carbonic acid of the air with the lime of the water. This carbonic acid gas supplies plants with one of their chief elements, carbon, for in the sunlight plants decompose the gas, retaining the carbon in their structures and liberating the oxygen.

The aqueous vapour present in the atmosphere is derived by evaporation from the sea and other surfaces of water. The amount in the air at a given place, depends on the situation of the place with regard to the sea, nature of the prevalent winds, &c., but varies considerably at different times. The presence of aqueous vapour in the air may be shown by cooling a vessel by means of a freezing mixture contained in it. The aqueous vapour in the air surrounding the vessel will be condensed and deposited on the sides of the vessel as dew or hoar frost.

Nitric acid and ammonia are found in the air combined—forming nitrate of ammonia. The substance is dissolved by the rain water, and, entering the roots of the plants, supplies them with nitrogen, an element which is needful for their growth, but which they cannot obtain directly from the air.

Ozone is found in air away from towns, and especially after thunder storms.

A cubic foot of air, at a temperature of 60° F. and 30 inches Bar., weighs 536·96 grains: or a cubic metre of air, at 0° C. and 760 m.m. pressure, weighs 1·2991 kilogrammes. At 60° F. a given volume of air is 816 times lighter than the same bulk of water at the same temperature.

The air extends to a height of at least 45 miles, and the pressure of a column of air of this height upon a square inch at the level of the sea is nearly 15 lbs.: or, we can say the pressure is equal to a column of mercury 30 inches, or 760 millimetres high, or a column of water nearly 34 feet high. The **barometer** is the instrument employed for measuring the pressure of the atmosphere. If the density of the air were the same throughout as it is at the sea-level, the height of the atmosphere would then be only about 5 miles.

The waters of the sea, rivers, and lakes, hold air in solution; but this air is richer in oxygen than the atmosphere, because that gas is more soluble than nitrogen. The air dissolved in water contains about 35 parts oxygen in every 100 volumes. In this way provision is made for the respiration of fishes.

DIFFUSION OF GASES.

Gases possess the remarkable property of mechanically mixing with each other in spite of differences of density. When liquids, which do not act chemically upon each other, are mixed together, they usually arrange themselves in layers according to their densities; and if gases followed the same law, the atmosphere, instead of being one uniform mixture, would consist, to some extent, of layers of different gases, the densest

(carbonic acid) at the surface of the earth, and the lightest in the upper regions of the air. This property of gases may be shown thus:—Fill two bottles with carbonic acid and hydrogen respectively: insert the bottle containing the hydrogen over the other, and connect the two by a glass tube, about two or three inches long, inserted in perforated corks in the mouth of each bottle. In a few hours the gases will be found uniformly mixed in the two bottles, some of the heavy carbonic acid having ascended, and the light hydrogen having partly descended.

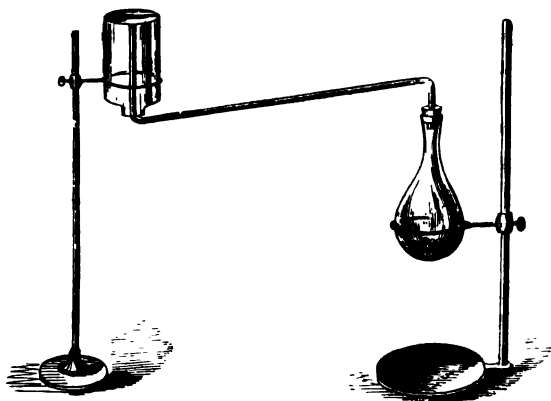
The rate at which they diffuse is not the same for all gases, since it has been found that a light gas diffuses more rapidly than a heavier one under the same conditions. The general rule has been established, that the velocity of diffusion of different gases is inversely proportional to the square root of their densities. For example, the density of oxygen is 16 times that of hydrogen, and the rate at which the diffusion of hydrogen takes place, compared with that of oxygen under the same conditions, is as 16 to 4, that is, as 4 to 1.

COMPOUND OF NITROGEN AND HYDROGEN.

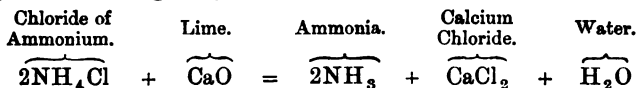
AMMONIA. *Symbol* NH_3 *Combining weight* 17.

This most important compound receives its name from the fact that one of its compounds (**sal ammoniac**, chloride of ammonium) was formerly prepared in Libya, near the temple of Jupiter Ammon. It occurs in small quantities in the air. It was prepared by the alchemists from fragments of horn, and hence it has received its common name of **spirits of hartshorn**.

Preparation. Ammonia may be prepared by heating in a closed vessel portions of any organic substance containing hydrogen and nitrogen. Most of the ammonia used in this country is obtained from the ammoniacal liquor produced in the manufacture of coal-gas. Hydrochloric acid is added to the liquor, and chloride of ammonium is thus formed. After evaporating to dryness, and further purification, the salt is ready for yielding ammonia.



Equal weights of quicklime and powdered chloride of ammonium are heated together in a retort, when ammonia comes off as a colourless gas, which may be collected by the displacement of air, or over mercury. One ounce of the salt yields about 750 cubic inches, or 30 grammes give about $12\frac{1}{2}$ litres of the gas. The change may be shown thus :—



Properties. Ammonia is a colourless transparent gas, with a strong, irritating smell, a bitter taste, and is feebly combustible. Its sp. gr. is 0.59. If collected by the displacement of air, the vessel must be inverted, and the tube conveying the gas inserted till it almost touches the top of the vessel. It neutralizes the most powerful acids, turns vegetable reds (*e.g.* litmus, which has been made red by an acid) blue, and in other respects behaves like an alkali. It forms an important class of salts.

If any volatile acid be brought into an atmosphere containing ammonia, a white cloud is produced, on account of the union of the acid and ammonia to form a solid salt. A glass rod, moistened with an acid, will in this way detect the presence of ammonia.

It is very soluble in water; water at 0° C. absorbs 1050 times its volume of ammonia, at 15° C. 727 times its volume. The solution has all the properties of the gas, and is of great use in the laboratory; it is colourless, and strongly alkaline. All the ammonia may be driven off from the solution by heat.

COMPOUNDS OF NITROGEN AND OXYGEN.

There are five compounds of Nitrogen and Oxygen, viz :—

1. Nitrogen Monoxide, or Nitrous Oxide (laughing gas)— N_2O .
2. Nitrogen Dioxide, or Nitric Oxide, N_2O_2 .
3. Nitrogen Trioxide, or Nitrous Anhydride, N_2O_3 .
4. Nitrogen Tetroxide, or Nitric Peroxide, N_2O_4 .
5. Nitrogen Pentoxide, or Nitric Anhydride, N_2O_5 .

These compounds serve to illustrate very clearly what is known in chemistry as the law of multiple proportions, which may be stated as follows:—When one body is capable of uniting with another in several proportions, these proportions bear a simple relation to each other. For example—28 parts by weight of nitrogen unite with 16 parts by weight of oxygen to form nitrous oxide; in nitric oxide, another compound formed by the combination of the same two elements, 28 parts of nitrogen combine with 32 ($= 16 \times 2$) parts of oxygen; in nitrous anhydride, 28 parts of nitrogen combine with 48 ($= 16 \times 3$) parts of oxygen; and so on with the remaining compounds. In the above cases, and in all other cases, the proportion by weight with which oxygen enters into combination is either 16, or some definite multiple of 16; of nitrogen, either 14, or some definite multiple of 14; of carbon, either 12, or some multiple of 12, &c.

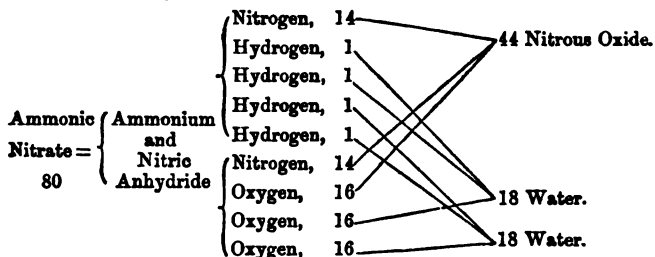
Three acids containing nitrogen, hydrogen, and oxygen, are known, viz :—

Hyponitrous Acid	- - - - -	HNO
Nitrous Acid	- - - - -	HNO_2
Nitric Acid	- - - - -	HNO_3

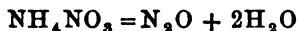
The last named of these acids is the only one which is of any practical importance.

NITROGEN MONOXIDE, Nitrous Oxide, or Protoxide of Nitrogen. *Symbol* N_2O . *Combining weight* 44.

This compound is prepared by heating ammonium nitrate (nitrate of ammonium) in a glass retort, when the salt is resolved into nitrous oxide and water. The gas may be collected at the pneumatic trough over warm water. The change may thus be shown (the numbers representing the combining weights) :—



OR IN SYMBOLS.

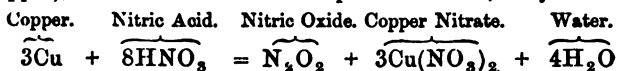


Properties. At ordinary temperatures this compound is a gas without colour and smell, but possessing a sweetish taste. Water at $15^\circ C$. dissolves about two-thirds its volume. It may be liquefied by the action of great cold or pressure. The sp. gr. of the gas compared with air is 1.527. It is, like oxygen, a brilliant supporter of combustion, but it is distinguished from oxygen by its solubility in water, its sweetish taste, and giving no ruddy fumes when mixed with nitric oxide. Its power of supporting combustion so brilliantly is due to the fact that it is readily dissociated by a burning body, and its elements, thus set free, form an atmosphere which is very much richer in oxygen than the air, containing over 36 per cent. of that gas. If respired for a few minutes it produces an intoxicating effect, hence it has the name of **laughing gas**; and on account of its also producing insensibility to pain, it is frequently employed in dental and surgical operations.

NITROGEN DIOXIDE, or Nitric Oxide. *Symbol* N_2O_2 (or NO). *Combining weight* 60 (or 30). *Density (compared with hydrogen)* 15.

This gas is prepared by pouring diluted nitric acid upon copper clippings in a retort; it may be collected over cold water. One ounce or 30 grms. of copper, acting upon 4 ounces of diluted nitric acid, yields nearly 420 cubic inches, or about 7 litres of nitric oxide.

During this decomposition, only one-fourth of the nitric acid is changed into nitric oxide. 2 out of every 8 parts of the acid give up the oxygen to 3 of copper, forming 3 atoms of protoxide of copper, 1 atom of water, and 2 of nitric oxide. The 3 atoms of the oxide of copper unite with the remaining 6 of nitric acid, and form 3 atoms of copper nitrate (nitrate of copper), and 3 of water are also separated. Thus, in symbols:—



Properties. It is a colourless, incombustible, and irrespirable gas, which resisted all attempts to liquefy it until Cailletet in 1877 converted it into a liquid at 11°C ., under a pressure of 104 atmospheres. It has a strong attraction for oxygen; if a jar of gas be exposed to the air, dense ruddy fumes (very soluble in water) are produced, in consequence of the union of the nitric oxide with the oxygen of the air, forming nitrogen tetroxide (N_2O_4). Ignited phosphorus will burn vigorously in this gas, provided its temperature when placed in it is sufficiently high to decompose the gas, and thus set oxygen free, otherwise the combustion is stopped. The same is the case with other burning bodies placed in this gas. The specific gravity of the gas compared with air is 1.039.

NITROGEN TRIOXIDE, or Nitrous Anhydride. *Symbol* N_2O_3 . *Combining weight* 76.

This is prepared by mixing 4 volumes of nitric oxide and 1 of oxygen in a flask, both in a dried state, and then cooling

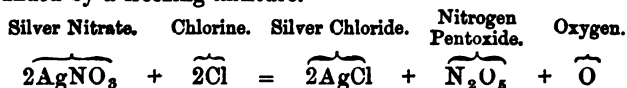
the mixture, which condenses into a blue volatile liquid. Nitrogen trioxide dissolves in cold water to form nitrous acid (HNO_2), hence it is also called **nitrous anhydride**. Nitrous acid forms salt with the metals of the alkalies, which are called **nitrites**.

NITROGEN TETROXIDE, Nitric Peroxide, or Hyponitric Acid. *Symbol* N_2O_4 (or NO_2). *Combining weight* 92 (or 46). *Density (compared with hydrogen)* 23.

This is prepared by mixing 2 volumes of nitric oxide with 1 of oxygen, or by heating dry nitrate of lead in a retort. The dense red fumes which nitric oxide forms when exposed to the air are chiefly the vapour of nitric peroxide. It can be condensed by cold into a volatile yellow liquid, or into a colourless crystallized solid.

NITROGEN PENTOXIDE, or Nitric Anhydride. *Symbol* N_2O_5 . *Combining weight* 108.

This compound may be prepared by passing dry chlorine gas over heated silver nitrate, and condensing it in a receiver surrounded by a freezing mixture.



The colourless crystals formed in this experiment unite energetically with water to form nitric acid.

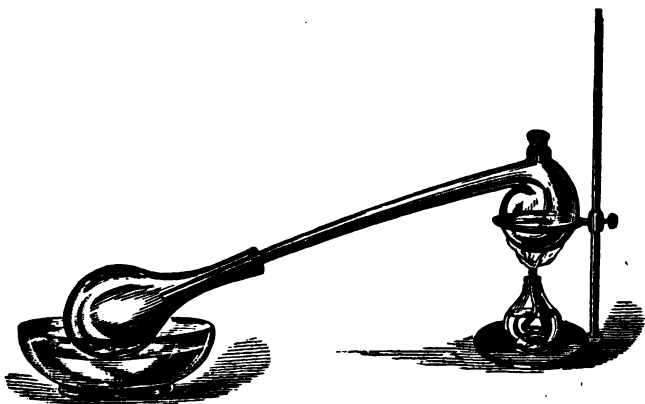
NITRIC ACID: or Hydrogen Nitrate. *Symbol* HNO_3 . *Combining weight* 63.

This is the most important compound of Nitrogen containing Hydrogen and Oxygen. It has long been known under the name of **Aquafortis**. Its true composition was determined in 1785.

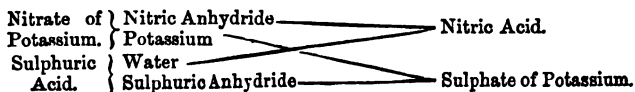
Nitric acid is found in the atmosphere, after thunder storms, in the form of nitrate of ammonia; also in tropical climates combined with the metals potassium and sodium. In various parts of India the soil is crusted over with nitrate of potassium,

or potassic nitrate, commonly called saltpetre or nitre. In the same way nitrate of sodium occurs in Chili and Peru, and is sometimes called Chili saltpetre.

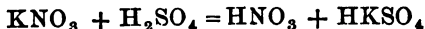
Preparation. Nitric acid may be formed directly by the combination of the elements nitrogen and oxygen, in the presence of moisture, when a stream of electric sparks is passed through the mixture. In practice it is always obtained from one of the nitrates. Place equal weights of nitrate of potassium (KNO_3) and sulphuric acid (H_2SO_4) in a retort. On applying heat, the potassium unites with the sulphuric acid to form



sulphate of potassium, which remains in the retort, while the nitric acid distils over into the receiver. Thus—



OR IN SYMBOLS.



Saltpetre and Sulphuric Acid give Nitric Acid and Sulphate of Potassium.

Properties. Nitric acid, when pure, is colourless, but it is usually of a yellowish colour, on account of the presence of some of the lower oxides of nitrogen. Its sp. gr. at 60°F . (15°C .) is 1.52 : it boils at 120.5°C .

Nitric acid is very corrosive, and dissolves nearly all the metals except platinum and gold. It very readily parts with a portion of the oxygen which it contains, and is therefore a powerful oxidising agent. It destroys all animal textures, and stains the skin, wool, and feathers of a bright yellow colour. If copper, zinc, tin, or iron filings be thrown into nitric acid, ruddy fumes (nitric peroxide) are given off, and the metals combine with the acid, and form salts called nitrates.

An acid is a compound containing hydrogen, the whole or part of that element being capable of replacement by a metal, as in the above instances. Some chemists regard an acid as a salt of hydrogen, and speak of nitric acid, for example, as "hydrogen nitrate."

Most of the salts called nitrates are soluble in water, fuse readily when heated, and at a high temperature are decomposed.

QUESTIONS ON THE ATMOSPHERE.

1. One hundred volumes of Nitrogen are given. How many volumes of Oxygen must be added in order to obtain a mixture of the same composition as pure atmospheric air? (L.)

2. State the constitution of the atmosphere, and shew how the proportions of its constituents may be determined. How is it affected by fires, and by the action of animal life? What is the natural compensation for this? (U.)

3. Explain how you would prove the pressure of Water, Carbonic Acid, Oxygen, and Nitrogen in atmospheric air. (L.)

4. What is the average composition of the atmosphere? Why is it considered a mechanical mixture and not a chemical compound? What several causes tend to keep its composition constant? (C.)

5. How would you show that air is necessary for the maintenance of ordinary combustion? How does air act on a burning substance? How could you show that the product of combustion of phosphorus in oxygen consists of phosphorus and oxygen? (L.)

QUESTIONS ON NITROGEN AND ITS COMPOUNDS.

1. Name the most important natural substances in which Nitrogen is contained; and describe how you would prepare Nitrogen. Enumerate the principal properties of Nitrogen, and show how it may be distinguished from Oxygen, Hydrogen, and Carbonic Acid. (L.)

2. Describe the mode of preparation of Nitric Acid, expressing in symbols the reactions which take place. Distinguish between a Nitrate and a Nitrite. (C.)

3. Describe the combinations of Oxygen with Nitrogen, and the processes by which they are respectively obtained. (O.)

4. What is the composition of Laughing Gas? How is it prepared? (C.)

5. What is the result of mixing over water Nitric Oxide and Oxygen? (O.)

6. Describe in words, and represent in symbols, the effects of heat upon Nitrate of Ammonia. Give some account of the chemical properties of the gaseous product. (L.)

7. Ten volumes of Nitrogen Protoxide are mixed in an eudiometer with 12 volumes of Hydrogen, and a spark is passed through the mixture; how many volumes of gas remain, temperature and pressure being uniform? (O.)

8. Give an account of the action of Hydric Nitrate upon various metals. (O.)

9. When Ammonium Chloride is heated with lime, a gas is given off. Explain how you would collect this gas, and describe some experiments to show its properties. (O.)

10. Explain the terms *Acid* and *Oxide*. Describe the properties of Nitric Acid and Nitric Oxide, and state how you would prepare them. (O.)

11. Two oxides are given you, one an Acid, the other a Base. Explain how you would determine which of the two is the acid. (L.)

12. Describe how you would prepare Nitrous Oxide. How would you distinguish this gas from Oxygen? (L.)

CARBON.

Symbol C. Combining weight 12.

This is the first solid element we have to mention. It is never found free in the state of a liquid or a gas. It is found uncombined in the solid state, as diamond and graphite. Charcoal is another form of this element. Though these substances are all carbon, yet they differ very remarkably in their physical characters.

Carbon is found as a characteristic constituent of all animals and vegetables. It is found in the air, combined with oxygen, as carbonic acid. In combination with oxygen, and with lime

and magnesia, it forms a great part of the earth's crust, as limestone, chalk, marble, corals, shells, &c.

Diamond is crystallized carbon with a sp. gr. 3.33 to 3.55. Lavoisier discovered this substance to be pure carbon in 1775-6, by burning it in oxygen, and converting it all into carbonic acid. It is the hardest substance known; and, when cut, possesses a brilliant lustre, and power of strongly refracting light. It is found in a crystallized form in certain rocks and gravels. The most famous diamond mines are those of Golconda and Bundelcund in India, of Borneo and of Brazil; and of late years many diamonds have been found in South Africa. It cannot be melted at any heat; but, placed between the poles of a powerful voltaic battery, it becomes a black mass resembling coke.

Graphite, commonly called "black lead," or "plumbago," has a sp. gr. 2.15 to 2.35. It occurs in veins in the granitic and primitive rocks, and is found in Cumberland, the United States, Ceylon, and Siberia. It is generally found mixed with a little iron and clay. The best variety is used for lead-pencils. The coarsest varieties are used for coating iron, polishing grates, and making crucibles which are to be exposed to great heat. Graphite is formed artificially in the process of coal-gas manufacture, and also as crystalline plates in certain varieties of cast iron.

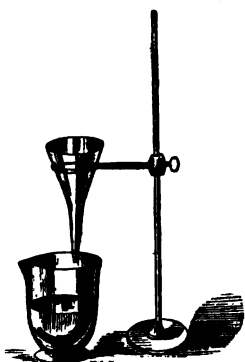
Charcoal is an artificial variety of carbon; so also are the varieties known as coke, lamp-black, and animal charcoal. Charcoal is formed by heating wood to redness in closed vessels, until all the water, oils, and gases are driven off. It is used in the manufacture of gunpowder, and in chemical experiments. It is very porous; it absorbs about ninety times its own volume of ammonia gas, and nine times its volume of oxygen. In consequence of its absorbent action it destroys bad odours, probably by causing the offensive gases condensed in its pores to combine with oxygen.

Coal consists for the most part of carbon, combined with

varying quantities of oxygen, hydrogen, nitrogen, and earthy matter. The variety of coal known as anthracite contains 90 per cent. of carbon.

Coke is the name given to the residue of coal, when treated in the same way as wood in the manufacture of charcoal. It is used as fuel, and especially for locomotive engines. It gives off little or no smoke, produces a steady and intense heat, and does not choke the draught by swelling or caking.

Lamp-black is the soot produced when coal-tar and similar substances are burnt with an insufficient supply of air. Mixed with linseed-oil and soap, it forms printers' ink. It is the purest form of artificial carbon.



Animal charcoal, or "bone charcoal," is obtained by heating bones in a retort. On account of its great decolourizing power, it is much used by sugar refiners to remove the colouring matter in raw sugar. This property may be seen by placing a little of it in water, coloured with indigo or litmus; boil the liquid, and then filter it, when a colourless liquid will be obtained.

Carbon is a very unalterable substance. Poles driven into the earth, and piles driven into the beds of rivers, are first charred to preserve them from decay.

COMPOUNDS OF CARBON AND OXYGEN.

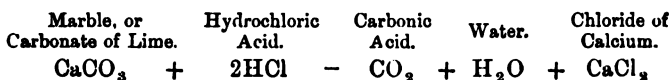
Carbon, when heated, unites with oxygen, forming two compounds—carbonic oxide (CO), and carbonic acid (CO_2).

CARBONIC ACID. (Carbonic Anhydride, or Carbon Dioxide.)
Symbol CO_2 : *Combining weight* 44. *Sp. gr. (referred to air)* 1.529.

This gas is found free in the atmosphere, in the proportion of about 4 parts by volume in 10,000 parts of air, in the water

of many mineral springs, and, combined with lime and magnesia, in limestone, corals, marble, shells, chalk, &c.

Preparation. It may be prepared in two ways, known as the analytic and synthetic processes. The synthetic mode is to burn carbon in oxygen, when each atom of carbon unites with two atoms of oxygen, and forms carbonic acid. It is obtained analytically by pouring some strong acid, as hydrochloric acid, on pieces of marble. Marble is a compound of lime (CaO) and carbonic acid, and gives off the latter when hydrochloric acid is poured upon it. The chemical changes may thus be shown :—



The gas may be collected by displacement of air (as shown in the accompanying illustration) on account of its great specific gravity, or it may be collected over water at the pneumatic trough.

Properties. It is a colourless, invisible gas, with a faintly acid smell and taste. Water dissolves about its own volume of

the gas; but the gas becomes more soluble if pressure be applied. When the pressure is removed the gas escapes with effervescence, as in soda-water, lemonade, champagne, &c. This gas is half as heavy again as air, and may be poured from one vessel into another. Under a pressure of 36 atmospheres at 0°C ., it is liquefied, and may be frozen into a snow-white solid. It cannot support combustion or animal life. A lighted candle put into a jar of the gas is immediately extinguished. It is injurious to health to breathe air containing over 6 parts of carbonic acid gas per 10,000 of air; and if the amount of carbonic acid gas present in the atmosphere were increased to from 3 to 4 per cent., animal life would probably be impossible. Hence the danger of living in crowded and ill-ventilated rooms.

It faintly reddens moist litmus paper; but when the gas has evaporated, the blue colour again returns. By combining with bases, it forms a numerous and important class of salts, called **carbonates**. The true acid, H_2CO_3 , is not known in a separate condition.

Natural Sources of Carbonic Acid. (1.) It is a product of animal respiration. If air from the lungs be sent through lime water, the latter becomes milky, on account of the formation of chalk (carbonate of lime).

(2.) It is produced during the process of fermentation. In breweries many accidents have happened by people descending into the vats, and becoming suffocated by the carbonic acid gas.

(3.) It is a product of burning limestone in lime-kilns.

(4.) In volcanic districts it often escapes from openings in the earth. The Dog's Grotto (Grotto del Cane), near Naples, is an example of this, the air in the grotto being said to contain about 70 per cent. of carbonic acid gas.

(5.) It is a product of the combustion of substances containing carbon in air or oxygen. Thus coal, gas, tallow, &c., when burning, give off this gas abundantly. Two tallow candles, when burning, produce as much carbonic acid as a man gives off by respiration in the same interval of time.

(6.) It is found in coal-mines after explosions of **fire-damp**, and is called by miners **choke-damp**. It accumulates in wells and pits; hence it is necessary to test the air by letting down a lighted candle before descending into one of these places.

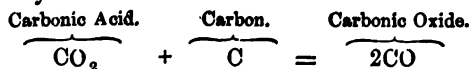
To test for the presence of free carbonic acid gas (in the air expired from the lungs, for example), use is made of a solution of lime in water (lime water). The gas is made to pass through the lime water, which is rendered turbid and milky by the formation of white carbonate of lime.

In the sunlight plants have the power of decomposing this compound into its elements, retaining the carbon to build up the woody parts of their structure, and setting free the oxygen. Animals, on the contrary, take in oxygen by respiration, which unites with carbon in their bodies to form carbonic acid gas. It will therefore be seen that animals and plants are complementary to each other as regards their chemical action on the atmosphere.

CARBONIC OXIDE: Carbon Monoxide. *Symbol* CO. *Combining weight* 28. *Sp. gr.* 0.967.

Preparation. This gas contains one atom less of oxygen than carbonic acid. It may be obtained by passing carbonic acid gas through a red-hot iron tube, when one atom of oxygen unites with the iron, and carbonic oxide comes out at the other end. In this way, one molecule of carbonic acid produces one of carbonic oxide.

It may also be produced by sending carbonic acid through a porcelain tube containing red-hot charcoal. In this case one molecule of carbonic acid produces two of carbonic oxide: for as the dioxide passes over the red-hot carbon, it parts with half its oxygen to the carbon, thus producing a molecule of carbonic oxide. In symbols—

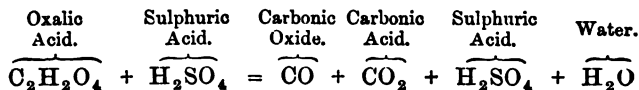


This latter process is constantly taking place in ordinary fires. The burning coals at the bottom of the grate produce



carbonic acid, which, passing through the red-hot parts in the centre, takes up carbon, and becomes carbonic oxide; this latter burns at the top of the fire with a blue flame, and changes again to carbonic acid.

For purposes of experiment, this gas may be conveniently prepared by heating together the deadly poison **oxalic acid**, ($\text{C}_2\text{H}_2\text{O}_4$) and strong **sulphuric acid**. The sulphuric acid takes the elements of water from the oxalic acid, which at once breaks up into the gases, carbonic acid and carbonic oxide. These come off together, but they may be separated by sending them through a solution of potash, which retains the carbonic acid. The decomposition may be shewn thus:—



Another good way of procuring this gas is to heat dry potassium ferrocyanide ($\text{K}_4\text{F}_2\text{C}_6\text{N}_6$) with sulphuric acid.

Properties. It is a transparent, colourless gas: sparingly soluble in water: it is very poisonous,—very small quantities, if taken into the lungs, cause death. The fatal effects of charcoal fires and limekilns are due to the presence of this gas.

It does not support combustion, but burns with a pale blue flame, producing carbonic acid.

QUESTIONS ON CARBON AND COMPOUNDS WITH OXYGEN.

1. What weight of Carbon is needed for the formation of 88 grammes of Carbonic Acid? What is the volume of this weight of acid at 0° C. and 30 inches bar.? (L.)

2. Describe how you would prepare Carbonic Acid from chalk; and explain your process by means of symbols. How would you test whether the air in a well consists of ordinary air or of Carbonic Acid? (L.)

3. How much carbon could be completely burned by a ton of air? (L.)

4. What volume of air is required for the combustion of one gramme of Carbon, the temperature being 15° C. and the height of the barometer 722 m.m.? (O.)

5. How many litres of Carbonic Acid (measured at 0° C. and 760 m.m. bar.) can be converted into Carbonic Oxide by 12 grammes of Carbon? and how many litres of Carbonic Oxide, at the same temperature and pressure, are thus produced? *

6. Eighteen cub. inches of Carbonic Oxide exploded, with 9 cub. inches of Oxygen, yield 18 cub. inches of Carbonic Acid. The specific gravities of Oxygen and Carbonic Acid are 1.1055 and 1.5197: find the sp. gr. of Carbonic Oxide. (O.)

7. A current of pure air passes through a bright coke fire. What gases are contained in the air beyond the fire? Give a short description of each of them. (L.)

8. If you place a burning candle in a closed bottle the flame will soon be extinguished. Why does the candle not continue to burn? How is the composition of the air affected by the flame? (L.)

9. Explain the action of Hydric Chloride (HCl) on Calcic Carbonate. If you pass Carbonic Acid over a layer of red-hot charcoal it becomes converted into a combustible gas. Explain the reaction which takes place between Carbonic Acid and hot charcoal. (L.)

10. The air of the Grotto del Cane is said to be composed of nearly pure Carbonic Acid. Describe experiments you would perform in order to test the accuracy of the statement. (L.)

COMPOUNDS OF CARBON AND HYDROGEN.

These compounds are very numerous, but their discussion belongs to the division of chemistry called organic. There are, however, two important compounds called light and heavy carburetted hydrogen, which must be considered here.

LIGHT CARBURETTED HYDROGEN. *Symbol* CH_4 . *Combining weight* 16. *Sp. gr.* 0.5576.

This gas is produced in stagnant pools by decaying vegetable matter, and is for that reason called **marsh gas**. It is given off from seams of coal in mines under the name of **fire damp**; and it is one of the chief constituents of coal gas.

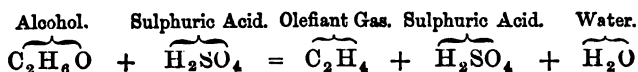
It may be obtained by heating together 40 parts of crystallized acetate of sodium, 40 parts of solid caustic potash, and 60 parts of powdered quicklime.

Properties. It is a colourless, tasteless, inodorous gas. Next to hydrogen, it is the lightest known substance. It burns with a yellow flame, forming carbonic acid and water. If mixed with twice its volume of oxygen, or ten times its volume of air, a violent explosion takes place on the application of a light. This is the mixture which causes explosions in coal-mines. The gases formed after the explosion, viz.—carbonic acid, nitrogen, and steam—are commonly named the “after damp,” or “choke damp.” Sir Humphry Davy and George Stephenson invented lamps to prevent these fatal explosions in coal-mines.

HEAVY CARBURETTED HYDROGEN. *Symbol* C_2H_4 . *Combining weight* 28. *Sp. gr.* 0.967.

This gas is an important constituent of coal-gas. It is known by the name of **Olefiant Gas**, because it forms an oily liquid when mixed with an equal volume of chlorine gas.

Preparation. Mix 1 measure of alcohol (spirits of wine) with 3 measures of sulphuric acid; heat the mixture in a retort, and olefiant gas comes off along with vapour of water. Thus:—



Properties. It is a transparent, colourless gas, having a sweetish taste. Under great pressure it may be liquefied. It

has a density nearly double that of Marsh Gas, hence it is called **heavy carburetted hydrogen**. It does not support life or combustion; but it burns with a white luminous flame.

When mixed with an equal volume of chlorine, oily drops are formed, hence the name of **olefant** (oil producing) gas. This oily body is also called **Dutch liquid**, because it was discovered in Holland. If mixed with 2 volumes of chlorine, and then ignited, it burns with much smoke, because its carbon is rejected, whilst its hydrogen unites with the chlorine to form hydrochloric acid.

When mixed with 3 volumes of oxygen, it explodes on the application of a light.

COAL-GAS. Coal-gas is a mixture of several gases, obtained by heating coal in closed iron retorts. The best coal for this purpose is **cannel coal**. When the coal is heated it gives off, in addition to the mixture known as coal gas, several substances, such as **tar, ammonia, water**. As these pass through a series of iron pipes, the tar, ammonia, and water are condensed and deposited, and the gas, after further purification, is received in large cylindrical iron vessels suspended in water.

Coal-Gas is purified in the following way:—As it issues from the retort, it passes through a large iron chamber half full of liquid into large perpendicular pipes, and there, in cooling, deposits the tar and ammoniacal liquor. The gas then passes into a chamber containing lime and hydrated ferric oxide, which remove the carbonic acid and sulphuretted hydrogen, and thence into the reservoirs or gas holders.

The removal of the sulphur compounds is very necessary, because, when burned, they produce sulphurous and sulphuric acids, which prove injurious to paintings and furniture.

The gas which passes into the receiver is a mixture of several gases. The following are usually present:—

Light carburetted hydrogen—about 40 to 50 per cent.

Hydrogen— ditto.

Olefiant gas— about 5 per cent.

Carbonic oxide.

Nitrogen.

Certain compounds of carbon and hydrogen, which give to coal-gas its peculiar odour, and increase its illuminating power.

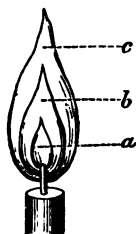
Sulphuretted hydrogen.

The proportion of the above substances depends upon the kind of coal employed. Olefiant gas is the chief illuminating substance. In comparing the light of coal-gas with that of candles, the gas from cannel coal, burning at the rate of five cub. ft. per hour, is said to be equal to 34·4 candles, burning 120 grains per hour, and common coal-gas equal to 13 candles.

THE NATURE AND STRUCTURE OF FLAME.

Flame is **gas** heated to such an extent that it becomes luminous. In order that flame may give off much light, it must contain solid matter, or dense vapours or gases, which become heated to whiteness. The amount of light does not depend solely upon the degree of heat, for the hottest flames, as the oxyhydrogen flame, give little light.

The flame of a candle, or lamp, has **three** distinct parts :—



(a) The dark central part (a) around the wick, containing the unburnt gases, called the **area of no combustion**.

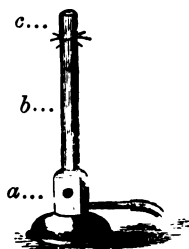
(b) The luminous part about (b), in which the hydrogen and carbon derived from the vaporisation and partial decomposition of the tallow, or other substance composing the candle, are only partially burnt, on account of the limited supply

of oxygen, called the **area of partial combustion**.

(c) The outside cone about (c), in which complete combustion takes place, giving off **great heat**, but little light. This is called the **area of complete combustion**.

In (*a*) the gases do not burn, because the oxygen of the air cannot reach that cone. In (*b*) the carbon in the solid state is heated to whiteness, because sufficient oxygen does not reach that part to burn it, hence the great light of this cone. In (*c*) the supply of oxygen is sufficient to burn the carbon, converting it into carbonic acid, and making the flame non-luminous.

By means of the **Bunsen** lamp, a flame of complete combustion may be obtained with ordinary coal gas. This lamp consists of an ordinary gas delivery-pipe, covered over with a tube about 4 inches long, with openings near the bottom. The air passes in at the openings (*a*), mixes with the gas in the tube (*b*), and burns on the top at (*c*) with a blue, smokeless flame. This lamp



is in general use in laboratories, on account of the great heat which it gives, its general convenience, and its cleanliness. Its cleanliness is due to the fact that there are no solid particles of carbon in the flame which can be deposited as soot on vessels in contact with the flame.

THE DAVY LAMP.

This lamp was invented by Sir Humphry Davy to protect miners from explosions in coal-mines. It consists of an oil lamp, enclosed in a cylinder of fine wire gauze, and furnished with the means of trimming the wick from the outside. The principle of the lamp depends upon the following facts :—Great heat is required to produce a flame. If the temperature of burning gas be lowered beyond a certain point, the flame goes out. A coil of cold copper wire, for example, placed over a flame, conducts away so much heat as to extinguish it. The flame, therefore, within the Davy lamp cannot pass through the wire gauze so long as it does not get red-hot, because the



metal reduces the temperature to such a point as to put out the flame in contact with it. But should the gauze by any means become red-hot, or should there be any break in it, the flame inside is communicated to any explosive gaseous mixture which may be outside.

This simple contrivance preserves mines from explosions of fire-damp. Marsh-gas, as we have seen, is combustible; but when mixed with a certain quantity of air, it becomes explosive. The gas in the pits passes through the wire gauze of the lamp, and burns within; but no flame can pass without to explode the mixture. Marsh-gas does not become explosive if mixed with less than 8 times, or more than 18 times, its volume of air.

THE BLOW-PIPE.

This is a small instrument for increasing the activity of combustion by throwing a jet of air into the interior of a flame.



The blow-pipe flame is divided into two parts, the outer (*b*) is called the **oxidising** flame, and the inner (*a*) the **reducing** flame. In the inner flame the

unburnt combustible gases, at a very high temperature, unite with the oxygen of any metallic oxide placed in it, and thus reduce the oxide to the metallic state, hence it is called the **reducing** flame. On the other hand, the heated air and gases at the point (*b*) of the outer flame, impart oxygen to a metal held there, and turn it into an oxide, hence it is called the **oxidising** flame. These properties of the blow-pipe flame may be seen by exposing a little lead on charcoal to the outer flame, when it becomes converted into the oxide of lead, and melts into a transparent globule. If this be now put into the inner flame, the oxygen is removed, and the lead is obtained in the metallic state.

The blow-pipe is a very useful instrument for the examination

of substances, with a view to ascertaining their chemical composition and other characters. Its use requires a continuous blast of air, which requires some practice to attain. The substance to be exposed to the blow-pipe flame is supported on a piece of charcoal, or on platinum foil, or by some other means.

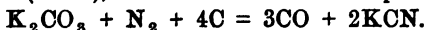


COMPOUNDS OF CARBON AND NITROGEN.

CYANOGEN. *Symbol* CN (or Cy.) *Combining weight* 26. *Sp. gr.* 1.806.

Cyanogen is so called from two Greek words meaning "blue producer," because it is the characteristic component of Prussian blue, and forms a number of other blue compounds.

Carbon and nitrogen do not combine directly; but if nitrogen gas be passed through a tube containing a mixture of charcoal and potassium carbonate heated to redness, the carbon and nitrogen unite with the potassium, forming a salt called potassium cyanide (KCN), while carbonic oxide escapes.



Potassium cyanide is the chief source of cyanogen and its compounds. This salt is obtained on the large scale by strongly heating, in a covered iron-pot, parings of horns, hides, hoofs, and other animal refuse, with potassium carbonate (pearl ash) and iron filings. The fused mass is then digested in boiling water, and the solution thus made yields on evaporation bright yellow crystals of potassium ferrocyanide, or yellow prussiate of potash, $\text{K}_4\text{FeC}_6\text{N}_6$. By heating this compound in a covered crucible, potassium cyanide, nitrogen, and a compound of carbon and iron are formed, and the potassium cyanide may be dissolved out of the residue by water. Potassium cyanide absorbs oxygen very readily, and is often used in blow-pipe analysis to reduce the metals from their oxides.

Preparation of Cyanogen. Dissolve in water potassium

cyanide and silver nitrate, in the proportion of 6·5 of the former to 17 of the latter ; silver cyanide will be formed as a precipitate. Dry the last named salt, and heat it in a test tube, and cyanogen will be given off as gas. Mercury cyanide may be used instead of silver cyanide.

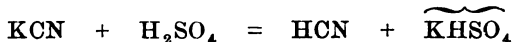
Properties. Cyanogen is a transparent, colourless gas, having a peculiar penetrating odour. It is very poisonous, combustible, and burns with a rose-edged purple flame, forming carbonic acid and free nitrogen. It condenses to a colourless liquid at a pressure of four atmospheres. It is soluble in water and alcohol.

Cyanogen combines with metals to form compounds called *cyanides*, e.g., potassium cyanide, KCN ; mercury cyanide, $\text{Hg}(\text{CN})_2$, &c.

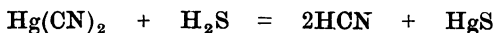
HYDROCYANIC ACID, OR HYDRIC CYANIDE (Prussic Acid).
Symbol HCN (or HCy). *Combining weight* 27. *Sp. gr.* 0·947.

Preparation. This important acid compound is obtained in a diluted state, by distilling a solution of the cyanide or the ferrocyanide of potassium with diluted sulphuric acid. The reaction may be shown thus :—

Hydrogen Potassium Sulphate.



Hydrocyanic acid may be prepared pure by passing sulphuretted hydrogen gas (H_2S) over dry mercury cyanide. The reaction may be thus represented :—



Properties. It is a colourless liquid, very volatile, boiling at $26\cdot5^\circ \text{C}$. The vapour is inflammable, and burns with a whiter flame than that of cyanogen. This acid is the most poisonous substance known, one drop being sufficient to cause death. Ammonia is one of the best antidotes. Its odour resembles that of the peach-blossom, or of bitter almonds.

Hydrocyanic acid is obtained by distillation from the kernels

of the bitter almond, and many kinds of stone fruits. It is present in the juice of the tapioca plant, and in water distilled off the leaves of the laurel, peach, and some other shrubs.

COMPOUNDS OF CARBON AND HYDROGEN, AND COAL-GAS.

1. How much pure air ought to be mixed with 100 measures of Fire-damp (Marsh-gas) in a coal-mine in order to supply material for the complete combustion of the Marsh-gas during explosion? (L.)
2. A cubic foot of Marsh-gas is mixed with 2 cubic feet of Oxygen, and exploded. What will be the resulting compounds, and what space will they occupy? (C.)
3. What chemical properties are characteristic of Olefiant gas? (O.)
4. What is the cause of an explosion in coal mines, and whence arises the danger of descending immediately after an explosion? (O.)
5. What volume of Olefiant gas must be used to form 10 cubic centimetres of Dutch liquid, of which the sp. gr. is 1.25? (O.)
6. Name the most important constituents, and some of the commonest impurities of Coal-gas. Give their chemical formulas, and explain what are the products to which they severally give rise when the gas is burnt. State which of the impurities may be removed, and by what means. (C.)
7. To what constituent does common Coal-gas chiefly owe its illuminating power? State how many volumes of Oxygen are necessary to its complete combustion, and name the resulting product. (C.)
8. How is Olefiant gas prepared? What is its composition, and how is this ascertained? Mention the distinguishing characteristics of this compound. (L.)

FLAME AND THE BLOW-PIPE.

1. Explain the nature of the chemical changes effected by the inner and outer flames of the Blow-pipe. (C.)
2. Describe the structure of the flame of a common candle, and explain why the flame loses its brightness when a current of common air is blown into it. (L.)
3. Explain fully why the flame of a common gas-light is brighter than the flame from a Bunsen burner. (L.)
4. What are the chemical and physical changes which take place when a candle burns? (O.)
5. A lighted candle is placed in a bottle, which is then securely stopped; what will take place, and what gases will be found in the bottle after some time? Explain how each of these gases comes to be there. (C.)
6. Explain the action of the Miners' Safety Lamp. (C.)

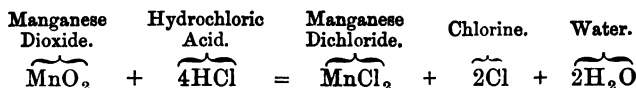
CHLORINE.

Symbol Cl. *Combining weight* 35.5. *Sp. gr. (referred to air)* 2.47.

The substances chlorine, bromine, iodine, and fluorine, are very much alike in their properties. They have great affinity for hydrogen, but very little for oxygen. They combine with hydrogen to form hydrogen-acids, consisting of one part of the element, and one part of hydrogen by volume. From these acids salts are formed, by the displacement of the hydrogen of the acid by a metal. These elements are sometimes spoken of as the *halogens* or "salt-producers."

Chlorine, the most important of the halogens, was discovered by Scheele in 1774. It is found most abundantly in nature, combined with sodium (common salt), and it also occurs combined with potassium and other metals.

Preparation. It may be obtained from any chloride. The most convenient way is to heat in a retort the black oxide of manganese (MnO_2), and hydrochloric acid (HCl), which is the chloride of hydrogen. The reaction is expressed as follows:—



The gas may be collected over warm water, because of its solubility in cold water, or it may be collected by displacement of air.

Properties. It is a gas of a greenish yellow colour, hence its name chlorine, from the Greek word $\chi\lambda\omega\rho\acute{o}\varsigma$ —green. It has a strong suffocating odour, but when present in small quantities it has the smell of sea-weed. It has a density about $2\frac{1}{2}$ times that of air. Under the pressure of 4 atmospheres, at 15.5°C ., it becomes a liquid. One volume of water at 15.5°C . dissolves about two volumes of chlorine, and the solution has the same odour as the gas, and a greenish tinge. This solution is a very useful laboratory agent, but it very soon loses its properties, for in the light the chlorine decomposes the water, uniting with the

hydrogen to form hydrochloric acid, and liberating the oxygen. The other halogen elements, when dissolved in water, undergo a similar change.

It has great affinity for hydrogen and metals, but little for oxygen and carbon. Powdered metallic antimony dropped into a jar of the gas takes fire spontaneously, forming antimonious chloride (SbCl_3). Other powdered metals and phosphorus will also burn when treated in the same way. The compounds formed by the union of chlorine with other elements are called **chlorides**. If a burning taper be plunged into a jar of chlorine, clouds of soot (carbon) are deposited, the chlorine uniting with the hydrogen of the hydrocarbons.

Chlorine, when associated with water, has strong bleaching power. Most of the vegetable colouring matters contain hydrogen, and since chlorine has great affinity for this gas, the colouring matter is decomposed. This property may be shown by the action of chlorine upon writing ink, litmus, the colouring matter of roses, &c. Printers' ink, which consists principally of lamp black (carbon), is not affected by chlorine. The gas in the free state, and in the form of bleaching powder, is used in cotton, linen, and paper manufactures.

Chlorine is also useful as a disinfectant, as it immediately decomposes all animal effluvia with which it comes in contact. It is used for this purpose in hospitals, &c., being usually set free from the compound known as "Chloride of Lime," or "bleaching powder."

HYDROCHLORIC ACID. *Symbol* HCl . *Combining weight* 36.5.

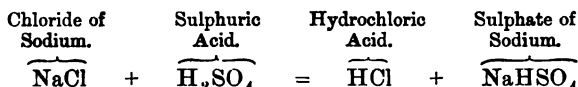
This most important compound of chlorine and hydrogen is also called **muriatic acid**, **spirits of salts**, and **hydric chloride** (chloride of hydrogen).

Preparation. It may be prepared in two ways—

(1). By mixing equal volumes of chlorine and hydrogen, and exposing them to sunlight, or applying a lighted match, when they will unite with explosion, and form hydrochloric acid.

(2). By heating in a retort common salt (chloride of sodium)

and sulphuric acid. Hydrochloric acid comes off as gas, which may be collected by displacement of air, or over mercury, and hydrogen-sodium sulphate (sulphate of sodium) is left behind in the retort. Thus :—



One ounce, or 30 grammes of common salt, yields 700 cubic inches, or 12 litres of the gas.

Properties. It is a colourless, invisible gas, with a density 1.269 times that of air, and possessing a most pungent odour. It is very soluble in water, forming the solution which commonly goes by the name of hydrochloric acid. One volume of water at 15.5° C. dissolves about 450 volumes of this gas. It may be liquefied at ordinary temperatures, under the pressure of 40 atmospheres. When allowed to escape into the air, it produces white fumes by combination with the moisture of the air. It is neither combustible nor a supporter of combustion. The strongest commercial hydrochloric acid contains about 43 per cent. of the gas, and has a specific gravity of 1.21. The commercial acid is generally impure; its yellow colour is owing to the presence of iron and other impurities. The great bulk of the hydrochloric acid made is used in the manufacture of bleaching powder.

The test usually employed for hydrochloric acid is a solution of nitrate of silver. This gives a white precipitate of chloride of silver (AgCl) when added to a solution containing free hydrochloric acid or a chloride. The precipitate is insoluble in nitric acid, but soluble in ammonia.

Two parts of hydrochloric acid, mixed with one of nitric acid, form aqua regia, which will dissolve gold and platinum.

COMPOUNDS OF CHLORINE AND OXYGEN.

These gases do not combine directly with each other, but compounds of them have been obtained indirectly. These com-

pounds are exceedingly unstable, a slight increase of temperature being usually sufficient of itself to decompose them, often with explosive violence. Three oxides of chlorine have been obtained in a separate state, viz:—

	Symbol.	Combining weight.
CHLORINE MONOXIDE (Hypochlorous Anhydride)	Cl_2O - - -	87
CHLORINE TRIOXIDE (Chlorous Anhydride) - - - - -	Cl_2O_3 - - -	119
CHLORINE TETROXIDE (Chloric Peroxide) - - - - -	Cl_2O_4 - - -	67.5

The following acids containing chlorine, hydrogen, and oxygen are known, the first two corresponding to the first two oxides, while the oxides corresponding to the remaining acids are not known:—

	Symbol.	Combining weight.
Hypochlorous Acid - - -	HClO - - -	52.5
Chlorous Acid - - -	HClO_2 - - -	68.5
Chloric Acid - - -	HClO_3 - - -	84.5
Perchloric Acid - - -	HClO_4 - - -	100.5

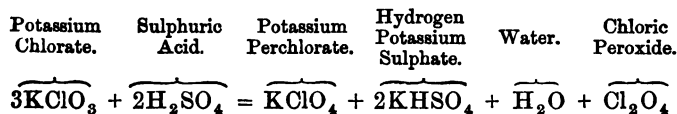
CHLORINE MONOXIDE, or Hypochlorous Anhydride (Cl_2O) is obtained by passing chlorine gas through a tube containing red oxide of mercury (HgO). It is a pale yellow gas, which may be condensed by cold to a deep red liquid. This vapour is very explosive when slightly heated. Combined with the elements of water it forms hypochlorous acid (HClO), which unites with bases to form salts, called hypochlorites.

BLEACHING POWDER, also known as "chloride of lime," is the best known hypochlorite. Its composition is somewhat ambiguous, as it is not simply a hypochlorite, but is regarded by some as a mixture of chloride and hypochlorite. It seems, however, to be best represented by the formula CaOCl_2 . It is prepared by passing chlorine gas over layers of moist slaked lime (calcium hydrate, CaH_2O_2), exposed on shelves in large

chambers. It is important on account of its bleaching and disinfecting properties, and large quantities of it are manufactured for the purpose of bleaching cotton goods. These properties it owes to the fact that its chlorine is very loosely combined, and readily set free by the action of weak acids, even the carbonic acid in the air is sufficient to liberate the gas gradually when the powder is exposed.

CHLORINE TRIOXIDE, or Chlorous Anhydride (Cl_2O_3) is a yellowish green gas, dangerous to prepare on account of its explosive properties. It dissolves in water, and forms chlorous acid (HClO_2)

CHLORINE TETROXIDE, or Chloric Peroxide (Cl_2O_4) is a greenish yellow gas, obtained by the action of sulphuric acid on chlorate of potassium.



By means of this gas phosphorus may be made to burn under water. Put a few pieces of phosphorus in a long narrow glass, add some chlorate of potassium, and cover them with water, then, by means of a long glass funnel, pour down a small quantity of strong sulphuric acid on to the chlorate, which will set free chloric peroxide, and the phosphorus will take fire. A mixture of powdered sugar and potassium chlorate takes fire for the same reason when a drop of strong sulphuric acid is allowed to fall on to it. No acid corresponding to this oxide is known.

CHLORIC ACID (HClO_3) is only of interest as combining with metallic oxides to form salts called **chlorates**. All chlorates, when heated, part with their oxygen (*e.g.*, the chlorate of potassium in the preparation of oxygen), and they are all soluble in water. Chlorate of potassium ("chlorate of potash")

is largely used in the manufacture of lucifer matches. It may be prepared by passing chlorine gas into a hot solution of potassium hydrate (KHO), when the chloride and chlorate of potassium are formed, and remain in solution.



The solution is then concentrated by evaporation, when, on cooling, the chlorate crystallizes out, leaving the chloride in solution. In manufacturing the chloride on the large scale, calcium hydrate is used instead of the much more expensive potassium hydrate, and calcium chlorate is formed, which is decomposed by the addition of potassium chloride, with the formation of calcium chloride, and potassium chlorate. No oxide corresponding to this acid is known: it would have the formula Cl_2O_5 .

PERCHLORIC ACID (HClO_4) is obtained by heating chlorate of potassium until one-third its oxygen is driven off. The mass which remains is a chloride and a chlorate, and from these the perchlorate of potassium is formed. It is also produced by the oxidizing action of hot nitric acid upon potassium chlorate. If strong sulphuric acid be mixed with the dry perchlorate, perchloric acid comes off, and condenses in the receiver as a colourless fuming liquid. It is one of the most powerful oxidizing agents known; a drop allowed to fall on paper, wood, or charcoal, produces combustion with violent explosion. Perchloric acid forms salts called **perchlorates**, which are all decomposed by heat.

CHLORINE AND NITROGEN.

These two gases may be made to combine indirectly by sending chlorine gas through a solution of ammonia, when drops of an oily liquid are seen to form. This liquid is known as **chloride of nitrogen** (NCl_3). It explodes with fearful violence when touched, producing free nitrogen and chlorine.

The composition of this compound is not well known—there are reasons for believing that it contains hydrogen as one of its constituents.

Two compounds of chlorine and phosphorus are known; viz., phosphorus trichloride (PCl_3) and phosphorus pentachloride (PCl_5).

QUESTIONS ON CHLORINE AND ITS COMPOUNDS.

1. Describe how you would prepare Chlorine, and explain your process. Also give an outline of the chief chemical and physical properties of this element. (L.)

2. Explain the action of Chlorine as a bleaching agent. A litre of Hydrogen gas at 0°C . and 760 m.m. pressure weighs 0.08986 gramme. What is the weight of the same quantity of Chlorine gas under the same conditions? (C.)

3. Describe and explain the process of making Hydrochloric Acid. (C.)

4. Determine the proportions in which the substances should be employed in making Hydrochloric Acid, so that there may be no waste; having given the equivalents of Sodium, Sulphur, and Chlorine to be as 23: 16: 35.5. (C.)

5. What weight of Oxygen, Nitrogen, and Chlorine, respectively, will combine with one part by weight of Hydrogen? And what volume of each of these gases will combine with one volume of Hydrogen? (O.)

6. To a solution of Hydrochloric Acid an excess of Carbonate of Sodium was added; the liquid was evaporated, and the dry residue weighed. The weight of Carbonate of Sodium added was 1.5 grms., that of the residue 1.6507 grms. What weight of Hydrochloric Acid did the solution contain? (O.)

7. A piece of marble is placed in dilute Hydrochloric Acid; state the effects observed, and the compounds formed. If a piece of iron be used instead of marble, state the results. (C.)

8. If, in the case of question 7, 4 fluid ounces of the acid will dissolve 1000 grains of marble, what quantity of iron would you expect an equal quantity of the acid to dissolve; the equivalents of calcium and iron being 20 and 28 respectively? (C.)

9. What is Chloride of Lime? Explain how it is prepared, and how it is applied in bleaching cotton fabrics. Would you expect stains of lamp-black or of nitrate of silver to be removed by these means? Why so? (C.)

10. Describe the preparation of Chlorine, and explain your process.

What happens when dry litmus paper, an aqueous solution of litmus, a burning candle, and powder of copper, respectively, are introduced into Chlorine gas? (L.)

11. A gas is found by analysis to consist of equal volumes of Chlorine and Hydrogen. Describe fully how you would find by experiment whether the gas is a mechanical mixture or a chemical compound. (L.)

12. What processes and materials would you employ to obtain crystallized potassium chlorate? How much oxygen is there in 100 grammes of the salt? (O.)

BROMINE.

Symbol Br. Combining weight 80. Sp. gr. as a gas, 5.54 (compared with air): as liquid, 3.187.

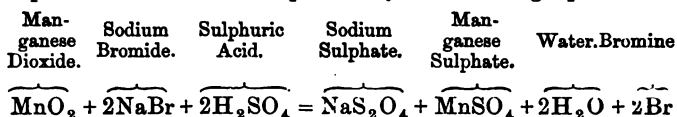
This element is much like chlorine in its properties and compounds. It was discovered by Balard, in 1826, in *bittern*,—the mother liquor of sea-water. It was named Bromine from *βρῶμος* (Bromos)—a **stench**, on account of its disagreeable smell.

It does not occur free in nature, but it is found combined with sodium and magnesium in sea-water and in the water of some mineral springs.

Preparation. Sea-water is evaporated until the more insoluble salts which it contains are removed by crystallization. The liquor is then called **mother liquor**, or **bittern**. Chlorine gas is sent through this, and liberates bromine from those metals with which it is combined, the chlorine taking its place. The bromine dissolves in the liquid, and gives it a yellow colour. Ether is then added to the solution, in order to dissolve the free bromine; and, on standing, the ether rises to the top of the liquid, forming a deep red layer.

Caustic potash, added to this solution, forms a bromide and a bromate of potassium. The ether may be distilled off, leaving the potassium salts. These latter are then strongly heated, when the bromate (KBrO_3) gives off all its oxygen, becoming converted into potassium bromide (KBr). This reaction

exactly corresponds with the reaction involved in the preparation of oxygen from potassium chlorate. The remaining bromide, mixed with black oxide of manganese and dilute sulphuric acid, and heated, gives off bromine as a deep red vapour. The reaction is expressed by the following equation:—



Properties.—Bromine is a deep red, heavy liquid: it is the only element, with the exception of mercury, which is liquid at ordinary temperatures. It has a strong and offensive smell, like chlorine, and is highly poisonous. It is slightly soluble in water, and the solution possesses bleaching powers. It is very readily soluble in ether, chloroform, and bisulphide of carbon. It boils at 63° C.; at 12° C. it becomes a red crystalline solid. It stains the skin yellow, like nitric acid. Free bromine may be recognized by its bleaching properties, and by the orange yellow stain which it gives to a solution of starch.

Like chlorine, it combines with hydrogen and forms **hydrobromic acid**, or hydric bromide (HBr), which is very soluble in water. Hydrobromic acid may be prepared by sending vapour of bromine and steam together through a red-hot tube. ($2\text{Br} + \text{H}_2\text{O} = 2\text{HBr} + \text{O}$). This compound is a colourless gas, and in its properties greatly resembles the corresponding chlorine compound, hydrochloric acid (HCl). Silver, potassium, and other metals may replace the hydrogen in HBr, forming compounds known as **bromides**.

There are also several oxides of bromine, of which **bromic acid** (HBrO_3) is the best known. **Bromates** (KBrO_3 , NaBrO_3 , &c.) may be formed directly from this acid by the displacement of its hydrogen by a metal.

Bromine combines directly with phosphorus and many of the metals, and forms bromides.

IODINE.

Symbol I. Combining weight 127. Sp. gr. of vapour, 8.716 (referred to air). Sp. gr. of solid, 4.95.

Iodine is found in sea-water, in combination with potassium and sodium, but in smaller quantities than bromine, and also in the water of some mineral springs. It was discovered by Courtois in 1812. It has also been found combined with silver in Mexico, and zinc in Silesia.

Preparation. It is obtained from kelp, the ash of certain sea-weeds. The kelp is digested in water, and the liquor reduced by evaporation to a small volume, the less soluble salts (NaCl, KCl, &c.) being removed as they crystallize out, until only the salts of iodine are left in solution. The iodine liquor is heated with black oxide of manganese and sulphuric acid, and iodine distils over as a deep violet-coloured vapour, which condenses to a bluish-black solid, with bright metallic lustre. The reaction is precisely similar to that which occurs in the preparation of bromine (see p. 62).

Properties. Iodine is usually met with in the form of a crystalline grey solid, generally as small plates or scales, having a metallic lustre. It melts at 107° C. and boils at 180° C., giving off a beautiful violet-coloured vapour, which condenses to grey-coloured crystalline scales. It smells something like chlorine; it stains the skin a brown colour. Taken internally it is poisonous; but it is used externally in medicine for glandular swellings. It is slightly soluble in water, to which it gives a yellow colour. It is readily soluble in ether and alcohol, giving brownish-red solutions, while it dissolves in carbon bisulphide, to form a violet-coloured solution. It is called Iodine (*ἰώδης*, violet-coloured), from the colour of its vapour.

It combines with many of the metals, forming iodides, which are decomposed by chlorine and bromine, the iodine being liberated. Iodine combines directly with phosphorus in the

cold, and the action is so energetic that the phosphorus bursts into flame.

Iodine forms with starch a beautiful blue compound. This property is made use of for the detection of small quantities of free iodine. A simple test for an iodide in solution is to add to the solution some starch paste, and then a drop of chlorine water. The chlorine liberates the iodine from its compound, and renders it free to act on the starch.

HYDRIODIC ACID (HI) is a compound of iodine and hydrogen, and may be prepared directly by passing iodine vapour and hydrogen together through a hot tube. It may also be obtained more conveniently by applying heat to a glass tube containing alternate layers of iodine, powdered glass, and chips of phosphorus. Hydriodic acid is a colourless gas, very soluble in water, and is in most respects very similar to hydrochloric acid.

A solution of hydriodic acid is best obtained by sending sulphuretted hydrogen through water containing iodine ($\text{H}_2\text{S} + \text{I}_2 = 2\text{HI} + \text{S}$). Hydrobromic acid may be prepared in a similar manner, by substituting bromine for iodine.

Certain metals unite with the iodine, liberating the hydrogen, and forming metallic iodides; *e.g.*, KI, ZnI_2 , &c.

COMPOUNDS OF IODINE AND OXYGEN.

Iodine has a greater attraction for oxygen than chlorine or bromine, and forms with it two acids, viz., **iodic acid** (HIO_3) and **periodic acid** (HIO_4).

Iodic acid may be prepared by boiling iodine in nitric acid. It combines with metals, and forms salts called **iodates**. These salts resemble the corresponding chlorates and bromates in giving off oxygen when heated, becoming converted into iodides.

Periodic acid is a substance of little interest.

COMPOUND OF IODINE AND NITROGEN.

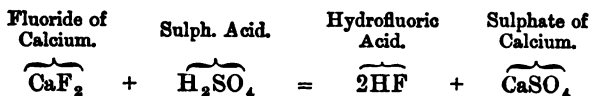
Iodine when dissolved in strong ammonia combines with the nitrogen, and forms a compound somewhat resembling the chloride of nitrogen. Like the latter substance, it is very explosive.

FLUORINE.

Symbol F. Combining weight 19.

This element is not found free in nature. It is found combined with the metal calcium, as fluoride of calcium (CaF_2), commonly called **fluor-spar**. It also occurs in a mineral called **cryolite**, in combination with sodium and aluminium. Fluorine has not yet been satisfactorily isolated. When set free from its compounds, it is so active chemically that it attacks the material of almost any vessel in which it is attempted to enclose it. It forms no compound with oxygen. Its most important compound is:—

Hydrofluoric acid. This answers to the hydrogen compounds of chlorine, bromine, and iodine, and may be prepared in a similar manner. Strong sulphuric acid is poured upon about half its weight of fluor-spar in the state of fine powder; heat is applied, and hydrofluoric acid comes off as dense fumes, which may be condensed to a colourless liquid in a suitably cooled receiver.

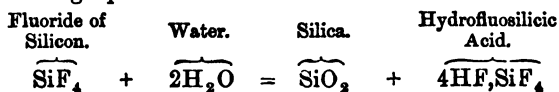


Hydrofluoric acid attacks glass, therefore it must be prepared in metallic vessels, platinum or lead being generally employed. At ordinary temperatures the liquid acid gives off dense fumes when exposed to the air, which are very corrosive, and care should be taken not to inhale them. The smallest

drop of the liquid acid upon the skin produces a deep and painful wound. Its most interesting property is that of etching glass, on account of the affinity of fluorine for the silicon contained in glass. This property may be shown by coating a piece of glass with wax, and, after removing a portion of the wax with a sharp point, exposing it to the fumes of the acid. The acid should be prepared as described above, in a vessel made by beating out a piece of sheet lead into the form of a cup, and covering it with the prepared glass plate.

When this acid is poured into water, it is dissolved, making a hissing noise. A weak solution of the acid is often employed, instead of the gas, for etching purposes.

Fluorine forms an interesting compound with silicon, called the **fluoride of silicon** (SiF_4). It is formed whenever free hydrofluoric acid comes in contact with silicon. The gas may be prepared by heating in a flask equal parts by weight of finely powdered fluor-spar and white sand, and eight parts of sulphuric acid, and collecting it over mercury. It is a colourless gas, decomposed by water. Each bubble of gas, as soon as it enters water, is changed into a bubble of gelatinous silica. The experiment is so arranged, that the delivery tube from the flask in which the gas is being prepared passes down to the bottom of a jar, and dips into mercury. Above the mercury in the jar is a quantity of water, which decomposes the bubbles of gas as they escape from the mercury. The object of the mercury is to prevent the mouth of the delivery tube from becoming stopped up with gelatinous silica. The reaction is expressed by the following equation :—



QUESTIONS ON IODINE, BROMINE, FLUORINE.

1. Describe Iodine. (O.)
2. What is the principal source from which Iodine is obtained? Describe and explain the process for extracting it, supposing it to be in the state of Iodide of Sodium. (C.)

3. Whence is Bromine obtained, and by what process ? (O.)
4. State the composition and the mode of preparing Hydrofluoric Acid.*
5. The Iodine in 100 volumes of Hydriodic Acid is liberated in succession by Chlorine and by Oxygen. How many volumes of Chlorine, and how many volumes of Oxygen are required ? *
6. One cubic foot of Hydriodic Acid is decomposed by an excess of Bromine, how many cubic feet of Hydrobromic Acid are formed ? *
7. A white salt is known to be either Chloride, Bromide, or Iodide of Potassium. How would you ascertain with certainty which of these substances it is ? (L.)

SULPHUR.

Symbol S. Combining weight 32. Sp. gr. of vapour at 500° C., 6.65; at 1000° C., 2.21, compared with air.

Sulphur is found in nature both free and combined. It occurs in the free state in beds of clay, &c., in Sicily, Iceland, and volcanic districts generally, and also in the form of powder on the sides of volcanic fissures. It is found combined with many metals as sulphides: thus—iron, copper, lead, zinc, are commonly found in combination with it. The most important natural sulphides are:—iron pyrites, copper pyrites, zinc blende, and galena (sulphide of lead). It is also found in combination with various bases in the state of sulphuric acid; these compounds are called sulphates, as the sulphate of calcium (**gypsum**), sulphate of barium (**heavy spar**, barytes), sulphate of magnesium (**Epsom Salts**).

Sulphur is obtained pure by heating a sulphide, as sulphide of iron (iron pyrites), in earthenware pots; the sulphur distils over in the form of vapour, and is condensed.

It occurs in commerce as **roll sulphur**, or **brimstone**, formed by casting melted sulphur into the form of sticks in moulds; **flowers of sulphur**, a crystalline powder formed by quickly condensing the vapour of sulphur; and **milk of sulphur**, a greenish white powder used in medicine, and precipitated by hydrochloric and other acids from certain alkaline sulphides.

Properties. Sulphur, at ordinary temperatures, is a yellow solid which is highly inflammable, burning in the air with a blue flame, and combining with oxygen to form sulphur dioxide, or sulphurous acid (SO_2), well known from its suffocating odour. It is insoluble in water, but is very readily soluble in bisulphide of carbon. It is a bad conductor of heat,—hence, when held in the hand, it crackles and splits.

It melts at 114.5°C ., and forms a thin yellow liquid, which becomes darker as the heat is increased. At about 250° — 260°C . it is so viscid that the vessel containing it may be quickly inverted without it running out; but at a still higher temperature it becomes quite liquid again. It boils at 445°C .

Sulphur can be obtained in three distinct modifications:—
(1.) **Octohedral:** native sulphur occurs frequently in beautiful rhombic octohedral crystals. When sulphur has been dissolved in carbon disulphide, and the latter has evaporated, the sulphur remains in the octohedral form. This variety is soluble in carbon disulphide, and has a sp. gravity of 2.05. (2.) **Prismatic:** when sulphur is melted in a crucible, and afterwards cooled so far that a crust has formed over its surface, if two holes be made in the crust, and the portion which is still liquid poured out, the crucible will be found filled with beautiful prismatic needle-shaped crystals of sulphur. This variety slowly passes into the octohedral, and is instantly transformed into that form when placed in carbon disulphide; the sp. gravity is 1.98. (3.) **Plastic:** when viscid sulphur is poured into water, the threads are at first elastic, insoluble in carbon disulphide, and have a sp. gravity of 1.95. The threads gradually become brittle, and the sulphur passes into the first variety which seems to be the most stable.

Sulphur is largely used in the arts. It is one of the components of gunpowder. It is also used in medicine. In the form of sulphurous acid it is used in bleaching silks and flannels, &c.

COMPOUNDS OF SULPHUR AND OXYGEN.

Sulphur forms two important compounds with oxygen, viz:—

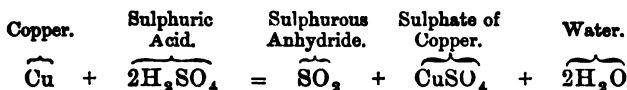
Sulphur Dioxide (Sulphurous Anhydride) - - SO_2

Sulphur Trioxide (Sulphuric Anhydride) - - SO_3

There are two acids corresponding to these two oxides, viz:—
sulphurous acid, H_2SO_3 , and sulphuric acid, H_2SO_4 .

SULPHUR DIOXIDE, or Sulphurous Anhydride. *Symbol* SO_2 ,
Combining weight 64. *Sp. gr., compared with air*, 2·247.

When sulphur burns in air or oxygen, this oxide is produced, which is known from its peculiar pungent and suffocating smell. This gas is frequently spoken of as "sulphurous acid," but that name strictly belongs to the compound H_2SO_3 . It issues in large quantities from the craters of volcanoes. When it is required pure, it is obtained by depriving sulphuric acid of one atom of oxygen. This is done on the small scale by heating copper with strong sulphuric acid. Thus:—



Properties. It is a colourless, invisible gas, very soluble in water. Water dissolves at 0°C . 68·8 times its volume of the gas: at 15°C . 43·5 times its bulk. On account of its great solubility in water, the gas prepared in the above experiment must be collected at the pneumatic trough over mercury, instead of over water. It may also be collected by displacement, on account of its great specific gravity, which is more than twice that of air. Sulphurous anhydride may readily be condensed to a colourless liquid, by surrounding the receiver with a freezing mixture, composed of pounded ice and salt. It may also be solidified. It is incombustible, and a non-supporter of combustion and respiration.

It possesses bleaching powers, and is used in bleaching straw, wool, silk, isinglass, sponge. Its action differs from that of

chlorine in not destroying the colour, which may be restored by neutralising the acid with an alkali, or expelling it by means of a stronger acid. It has also disinfectant properties.

SULPHUROUS ACID (H_2SO_3) is a somewhat unstable compound, formed when SO_2 is dissolved in water. Combined with metals it forms salts called *sulphites*. It gradually absorbs oxygen from the air, and becomes converted into sulphuric acid. This affinity of sulphurous acid for oxygen is utilised in the manufacture of sulphuric acid, as described below.

SULPHUR TRIOXIDE, or Sulphuric Anhydride. *Symbol* SO_3 . *Combining weight* 80.

This oxide is a white crystalline or fibrous solid, which may be prepared directly by passing sulphur dioxide and oxygen together over finely divided platinum, heated to redness. It may be also prepared by heating the strongest or Nordhausen sulphuric acid. It has an intense affinity for water, when thrown into that liquid it unites with it with a hissing noise, producing sulphuric acid.

SULPHURIC ACID. *Symbol* H_2SO_4 . *Combining weight* 98.

This, the most important of all the acids, rarely occurs in nature free, but it has been met with in the water of springs in volcanic districts. It occurs abundantly in combination with calcium, magnesium, and barium, in the form of sulphates. It is manufactured in England to a large extent, especially in South Lancashire; and more than 100,000 tons are consumed annually in this country.

Preparation. Ordinary sulphuric acid is prepared by uniting one atom of oxygen to sulphurous acid, by means of nitric acid and steam. This process is as follows:—Sulphur (or iron pyrites) is burnt in a furnace, and forms sulphurous anhydride. This is conveyed into large leaden chambers, along with nitric acid and steam. The nitric acid parts with a portion of its

oxygen to the sulphurous anhydride, and the sulphuric acid thus formed condenses on the floor of the chamber in combination with water, and is drawn off.

The nitric acid, on parting with its oxygen, is changed to nitric oxide (N_2O_2), and this, as soon as formed, combines with the oxygen of the air freely admitted into the chamber, and becomes nitric peroxide (N_2O_4). This nitric peroxide again parts with its oxygen to the sulphurous acid, and returns to the state of nitric oxide. Thus these compounds of nitrogen act as carriers of oxygen to sulphurous acid, and a very small quantity of them is capable of oxidising large quantities of sulphurous acid.

As the sulphuric acid, when first drawn off, is too weak, it is evaporated in leaden pans until it has the sp. gr. of 1.848. It then forms the oil of vitriol of commerce.

In Nordhausen, in Saxony, sulphuric acid is prepared by heating the sulphate of iron (green vitriol). This acid goes by the name of Nordhausen sulphuric acid, and is a brown fuming liquid, more concentrated than ordinary sulphuric acid, having a sp. gr. of 1.9.

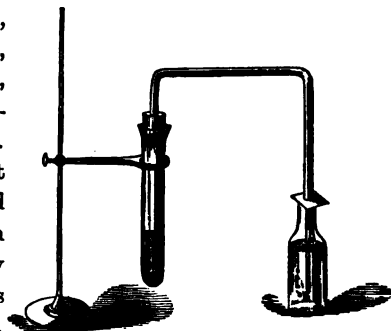
Properties. The sulphuric acid of commerce, commonly called oil of vitriol, is a dense, colourless, oily liquid, boiling at $325^{\circ}C$. It has a great affinity for water, and when mixing with it, gives off great heat. Great care should therefore be used in diluting this acid; the acid should be poured into the water gradually, and the mixture constantly stirred. This property of the acid is made use of in drying substances that cannot be exposed to heat. It has a characteristic sour taste, and reddens vegetable-blues (litmus, &c.) It is corrosive, destroying all organic substances it touches, usually charring them by abstracting from them the elements of water.

It dissolves most of the metals cold or when heated. Zinc, iron, cobalt, nickel, manganese, are dissolved by dilute sulphuric acid, hydrogen being given off, and sulphates of the metals formed. It displaces most other acids from their combinations.

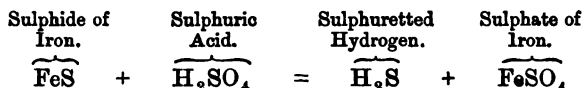
Several other oxy-acids containing sulphur are known, but are of little importance, viz:—Hyposulphurous acid, H_2SO_2 ; Thiosulphuric acid, $\text{H}_2\text{S}_2\text{O}_3$; &c. The oxides corresponding to these acids are unknown.

SULPHUR AND HYDROGEN.

HYDROSULPHURIC ACID, Sulphuretted Hydrogen, or Hydrogen Sulphide, (*Symbol* H_2S —*Combining weight* 34—*Sp. gr.* 1.191), is an important compound of sulphur and hydrogen. It occurs in combination with many metals, these compounds being known as sulphides, or occasionally as sulphurets.



Preparation. It is generally prepared by adding diluted sulphuric acid to the sulphide of iron (FeS). The iron and hydrogen change places (two atoms of hydrogen replacing one atom of iron); sulphuretted hydrogen comes off as a gas, and sulphate of iron or green vitriol remains. Thus:—



Properties. It is a colourless gas, having the smell of rotten eggs. It is very poisonous when respired; a small quantity is sufficient to destroy the lower animals. It burns with a pale bluish flame, producing sulphurous acid and water ($\text{H}_2\text{S} + 3\text{O} = \text{H}_2\text{O} + \text{SO}_2$). It is very soluble in water, which at 32° F. dissolves 4.37 its volume of the gas. It is less soluble in warm than in cold water, as is the case with gases generally; it is therefore sometimes collected over tepid water. The solution

possesses the properties of the gas, but it is gradually decomposed by the oxygen of the air, depositing its sulphur.

Sulphuretted hydrogen may be easily condensed by cold and pressure to a colourless liquid, and even to a solid, resembling ice.

The gas is immediately decomposed by chlorine, bromine, and iodine, the hydrogen combining with these substances, and the sulphur being set free, thus— $2\text{Cl} + \text{H}_2\text{S} = 2\text{HCl} + \text{S}$.

Sulphuretted hydrogen occurs in certain springs, as, for instance, in Harrogate waters. It frequently occurs as one product of the putrefaction of animal and vegetable substances.

Sulphuretted hydrogen is of constant use in the laboratory as a test for various metals. If this gas be added to acid solutions of arsenic, antimony, tin, gold, platinum, silver, mercury, lead, bismuth, copper, and cadmium, it immediately precipitates sulphides of these metals. If, however, it be added to acid solutions of iron, chromium, aluminium, cobalt, manganese, nickel, or zinc, no precipitates are formed, because the sulphides of these metals are soluble in acid solutions, whereas those first mentioned are insoluble; but they are precipitated on the addition of ammonia:—thus this gas serves to divide those metals into two groups. Most of the above precipitates are black. The precipitate in an arsenic or cadmium solution is **yellow**; with tin, it is **brown or yellow**; with antimony, **orange-coloured**; with zinc, **white**; with manganese, **flesh-coloured**.

As this gas is present in small quantities in coal-gas, when the gas has not been properly purified, the use of the latter is injurious to paintings, and tarnishes the silver goods of jewellers, &c., both from the H_2S itself and from the SO_2 formed by it on combustion.

A very simple test for the presence of H_2S in a gaseous mixture is to moisten a piece of white blotting paper with a solution of acetate of lead (sugar of lead), and to place it in contact with the gas. If H_2S is present, a black or brown stain

will be formed on the paper, from the production of sulphide of lead.

CARBONIC DISULPHIDE, or Bisulphide of Carbon. *Symbol* CS_2 . *Combining weight* 76. *Sp. gr. of liquid*, 1.27.

This compound is formed by passing vapour of sulphur over red-hot carbon. It is a colourless liquid, insoluble in water, but it is very soluble in ether and chloroform. It evaporates very rapidly when exposed to the air. Both the liquid and vapour are poisonous. It is used as a solvent for many oils and fats. Gums, caoutchouc, sulphur, iodine, bromine, and phosphorus dissolve in it. It burns with a blue flame, producing sulphurous acid and carbonic acid.

It will be seen that this compound is analogous to carbonic acid, with the oxygen of that compound replaced by sulphur, just as H_2S is the analogue of H_2O . Salts corresponding to the carbonates are known, containing sulphur instead of oxygen, and called sulpho-carbonates; thus— CaCO_3 , calcium carbonate; CaCS_3 , calcium sulpho-carbonate.

QUESTIONS ON SULPHUR AND ITS COMPOUNDS.

1. Describe the chief allotropic modifications of Sulphur, and name the chief substances from which Sulphur is obtained. How is Sulphurous Acid obtained? How much Sulphur, and how much Oxygen, are contained in 100 grammes of Sulphurous Acid? (L.)
2. Describe the process of manufacturing Sulphuric Acid, and state its most important uses. (O.)
3. By what different methods can the oxidation of Sulphurous Acid be effected? Explain the principles on which they severally depend? (O.)
4. Describe the method of obtaining Sulphuretted Hydrogen. For what purposes is it employed? Mention some cases in which it is naturally produced. What means would you employ to remove it from a tainted atmosphere, and how does the substance you name act upon it chemically? (C.)
5. You have given to you iron filings, Sulphur, and Hydrochloric Acid, and are required to make Sulphuretted Hydrogen, how will you proceed? *

6. What compound is produced when Sulphur is burnt? How may this be converted into Sulphuric Acid? (C.)
7. How much Oxygen, Sulphur, Hydrogen, and Nitrogen are contained in 100 grains of H_4NOSO_3 ? (C.)
8. Write a short account of the element Sulphur. (O.)
9. What compounds are formed when Sulphuretted Hydrogen is burnt in air; when it is mixed with Silver Nitrate; and when it acts on Mercuric Oxide respectively? (C.)
10. Express by equations the reactions which occur when Sulphuric Acid is heated together with Common Salt, with Sal-ammoniac, with Nitre, and with Copper respectively. (L.)

SELENIUM.

Symbol Se. *Combining weight* 78. *Sp. gr. of solid*, 4.78.

This is a rare element, discovered by Berzelius, in 1817, and called by him Selenium, from the Greek word $\sigma\epsilon\lambda\eta\nu\eta$ (Selene), the moon. It has properties much like sulphur, and is always found in combination with metals, as selenides, which are associated in nature with sulphides.

It may be obtained in three states, viz., amorphous, vitreous, and crystalline. At ordinary temperatures it is a solid of a deep brown colour, and somewhat metallic lustre. It has neither taste nor smell, and is insoluble in water, but soluble in carbon disulphide. When heated in closed vessels, it gives off a deep yellow vapour. It burns in air with a bright blue flame, giving off an odour resembling that of horse-radish.

With oxygen, selenium forms two compounds, which, combined with water, are acids, and correspond to the sulphur compounds, sulphurous and sulphuric acid.

SELENIOUS ANHYDRIDE (SeO_2) is formed when selenium is burnt in air or oxygen. When combined with one molecule of water it becomes selenious acid (H_2SeO_3). With bases it combines to form selenites.

SELENIC ACID (H_2SeO_4), corresponding to sulphuric acid, has not been obtained in the anhydrous condition.

Selenic acid has properties resembling sulphuric acid, and its salts, called seleniates, correspond to the sulphates. It is the only single acid which dissolves gold. On platinum it has no action.

With hydrogen, selenium forms a compound—seleniuretted hydrogen (H_2Se), which may be obtained by the action of an acid upon a selenide. It has a most disagreeable smell, and if respired in the smallest quantity, it causes great irritation.

TELLURIUM.

Symbol Te. *Combining weight* 128.

This element is very closely allied, in its chemical characters, to sulphur and selenium, but in most of its physical characters it is allied to the metals. It is met with very rarely in nature, but occurs as a constituent of certain gold and silver minerals, found in Transylvania and Hungary. It has a white colour, a metallic lustre, and a specific gravity of 6.2 compared with water.

Its compounds are analagous to those of sulphur and selenium, for example—Telluretted hydrogen, H_2Te ; Tellurium dioxide, TeO_2 ; Telluric acid, H_2TeO_4 .

SILICON.

Symbol Si. *Combining weight* 28. *Sp. gr.* 2.49.

Next to oxygen this is the most abundant element in nature. It is always found combined with oxygen, forming silica. This compound forms the stony skeleton of the earth. Rock-crystal, sand, and flint consist of silica in a more or less pure state, and it forms the chief constituent of many other minerals. It receives its name from flint or sillex.

Silicon may be obtained in three states, corresponding to the three states of carbon, viz :—

(1.) As a brown amorphous powder, by heating the silico-fluoride of potassium ($4KF, SiF_4$) with potassium. This form of silicon burns brilliantly in air, and is changed into silica.

(2.) In the graphitoidal form, like the graphite form of carbon, by heating the amorphous powder to a high temperature.

(3.) As a crystalline body, by fusing the above mixture, which gives amorphous silicon, with zinc. Crystals of silicon are formed on the zinc, and the latter may be removed by solution in an acid. The crystals are hard enough to cut glass.

SILICIC ACID (SiO_2) or **Silica**, is the only known oxide of silicon. It occurs pure in **rock-crystal**, **white sand**, and some varieties of **quartz**. It is found less pure in **sandstone**, **flint**, and a great variety of minerals. It enters largely into the composition of many precious stones, as **agate**, **amethyst**, **onyx**, **opal**.

Silica may be obtained in the form of a powder, as follows :— One part of finely divided quartz or white sand is heated with four parts of carbonate of soda ; the mixture fuses, and gives off carbonic acid ; the mass is boiled with water, and dissolves, when, on the addition of hydrochloric acid, hydrated silicic acid (H_4SiO_4) separates as a gelatinous mass. If this be evaporated to dryness, then heated, and hydrochloric acid added, silicic acid is obtained as a white powder. Another method of obtaining silica, in a state of great purity, is by the action of water upon fluoride of silicon (see p. 66).

Properties. Silica is a white, tasteless powder, insoluble in water and acids (except hydrofluoric acid) ; but it dissolves in strong alkaline solutions. Its sp. gr. in the crystallized form is 2.64, in the amorphous form 2.2. It can only be fused by the oxyhydrogen flame.

In the gelatinous form it is slightly soluble in water, but the solution has no action upon vegetable colours. It combines with bases, and forms compounds called **silicates**. The natural

silicates are among the most common of minerals, *e.g.*, clay, mica, felspar, &c. Nearly all the silicates are insoluble in water, but the silicates of sodium and potassium are mostly soluble. Silica enters largely into the composition of glass, porcelain, and earthenware, which are mixtures of various silicates (aluminium, lead, potash, &c.).

Silicon forms compounds with hydrogen, sulphur, chlorine, and fluorine. Compounds of the latter have been described under fluorine; the others are of little importance.

BORON.

Symbol B. Combining weight 11. Sp. gr. 2.68.

Boron resembles carbon and silicon in many of its properties. It is a body sparingly found in nature, and it is always found combined with oxygen as **boracic acid**. From this acid boron may be obtained in **two** forms,—crystalline and amorphous. The amorphous form is obtained by heating boracic acid with potassium or sodium. It is a dark greenish brown powder, which burns in the air when strongly heated, and burns with great brilliancy in pure oxygen, in both cases producing boron trioxide (B_2O_3). The crystalline variety is produced by heating the amorphous form with aluminium; the crystals are nearly equal to the diamond in lustre and hardness.

Boron combines directly with nitrogen, burning when heated in that gas, and producing **boron nitride** (BN). This is one of the very few cases in which nitrogen unites directly with another element.

Boron unites also with chlorine and with fluorine, forming **boron trichloride** (BCl_3), and **boron trifluoride** (BF_3). Boron is a triad element, as evidenced by these two compounds.

BORACIC ANHYDRIDE (B_2O_3). *Combining weight 70.*

This is the only oxide of boron. It is found free in the

lagoons of Tuscany, issuing along with steam from vents in the earth in that volcanic district. It is also found combined with sodium as borax or tincal, in the waters of salt lakes, in Thibet and California. A borate of lime and magnesia comes from South America.

BORACIC ACID, or Boric Acid (H_3BO_3) is prepared by adding sulphuric acid to a hot solution of borax, when sulphate of sodium is formed, and boracic acid separates in the form of white pearly crystals. On heating, the crystals part with the water, and pass into a fused glassy mass.

Boracic acid is slightly soluble in water. Its taste is faintly sour; it turns yellow turmeric paper brown, and gives to litmus a reddish tint. It dissolves in alcohol, and the solution burns with a green flame.

BORAX, the baborate of sodium, is the most important compound of boracic acid. Its symbol is $Na_2B_4O_7 + 10H_2O$. This substance is useful in the manufacture of the finer kinds of glass, and in forming glazes for porcelain.

Fused borax has the property of dissolving many metallic oxides, and forming, during the process, coloured glasses. This property makes this substance very useful in the laboratory for detecting certain oxides. A small quantity of borax is taken upon a loop of platinum wire, and fused, so as to drive off its water of crystallization. A small quantity of the substance under examination is then placed on the fused borax, and the whole again exposed to the flame. The substance is dissolved by the borax, forming a glassy bead, which is usually coloured. From the colour of the bead, the exact nature of the substance under examination may, in many cases, be inferred. Thus, an oxide of chromium gives, with borax, a bead of green colour; an oxide of copper, one of red in the reducing flame; an oxide of manganese, one of violet (*see* Analytical Table, No. I., p.).

QUESTIONS ON SILICON AND BORON.

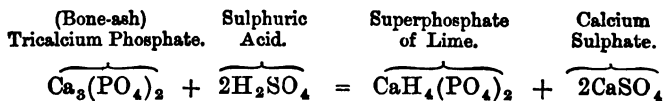
1. What are the sources and combinations of the element Boron? (O.)
2. How has Silicon been obtained? (O.)
3. What is Borax? Explain its use in chemical analysis. (C.)
4. Silica has no taste, and does not act on blue litmus. Why do we call this substance silicic acid? (L.)
5. A splinter of a colourless, transparent, solid body is given to you; how would you ascertain (a) by physical tests, (b) by chemical tests, whether this is glass or diamond? (L.)
6. A crystal is given to you, and you are requested to ascertain by a chemical experiment whether it is a Diamond or a common Rock-crystal. Describe an experiment which will furnish the required information. (L.)

PHOSPHORUS.

Symbol P. Combining weight 31. Sp. gr. of solid, 1.83.

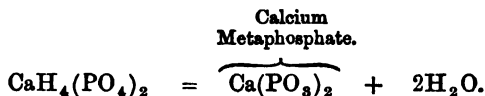
Phosphorus, which was discovered by Brandt of Hamburg in 1669, is not found in nature free; but in the state of phosphoric acid, combined with metals (calcium, lead, &c.), it occurs in primitive and volcanic rocks. From the soil formed by these rocks, plants extract phosphorus, and these latter supply the animal system with this substance. The bones, the brain, the nerves, and various tissues of animals, have phosphorus as one of their most important ingredients.

Preparation. The ashes of bones, which chiefly contain phosphate of calcium (lime), are acted upon by two-thirds of their weight of diluted sulphuric acid. The acid dissolves a part of the calcium, forming sulphate of calcium, which separates as an insoluble white powder. The remaining liquid contains in solution the phosphorus, combined with calcium and oxygen, as superphosphate of lime.



This liquid is filtered off and evaporated till it becomes as

thick as syrup, then mixed with charcoal and heated in an iron pot, till all the moisture is driven off, and the superphosphate of lime is converted into calcium metaphosphate.



The mixture is afterwards heated in an earthenware retort, the neck of which dips under water. The phosphoric acid (P_2O_5) of the superphosphate of calcium is decomposed by the charcoal; the oxygen unites with the charcoal to form carbonic oxide, and the phosphorus distils over, and condenses under water in yellow drops.



Properties. Phosphorus is a slightly yellow solid, much resembling wax, soft and flexible when warm, but brittle at temperatures near 0°C . It melts at 44°C . and boils at 290°C . It is very inflammable, and requires to be kept under water, and to be manipulated with great care, always under water. When exposed to the air it gives off white fumes, and these when seen in the dark are slightly luminous, emitting a pale greenish light. It receives its name phosphorus (from $\phi\omega\varsigma$ —light, and $\phi\acute{\epsilon}\rho\omega$ —I bear) from this property. It is not soluble in water, but dissolves in oils and bisulphide of carbon, and from these solutions it may be obtained in a crystallized form. Phosphorus is highly poisonous.

Phosphorus may be obtained in various forms, but the most interesting is red, or **amorphous** phosphorus. This is obtained by heating yellow phosphorus for many hours in a vessel containing carbonic acid, or other gas in which phosphorus is chemically inactive, at a temperature of from 230° — 250°C . The red kind differs very much in its properties from the yellow kind. It does not combine with the oxygen of the air at ordinary temperatures, and therefore need not be kept under

water, and it does not take fire until heated above 260°C , when it is converted into the ordinary yellow variety. It is not soluble in oils, bisulphide of carbon, &c., and its specific gravity is 2.14.

Red phosphorus is used in the manufacture of safety matches. The matches, after being dipped in melted paraffin, are tipped with a paste consisting of potassium chlorate, antimonious sulphide, powdered glass, and gum water. These matches ignite when rubbed upon a surface covered with a mixture of red phosphorus and powdered glass. Common lucifer matches are tipped with a mixture of phosphorus and potassic chlorate, and these take fire when rubbed on any rough surface.

COMPOUNDS OF PHOSPHORUS AND OXYGEN.

Phosphorous and oxygen form the two following compounds :—

	Combining Weight.
Phosphorus trioxide (Phosphorous Anhydride P_2O_3 ...	110
Phosphorus pentoxide (Phosphoric Anhydride) P_2O_5 ...	142

HYPHOPHOSPHOROUS ACID (H_3PO_2) has no corresponding anhydride. When phosphorus is boiled with a solution of baryta, water is decomposed, giving rise to phosphoretted hydrogen, phosphoric acid, and hypophosphorous acid; the first escapes as gas, whilst the last two remain in union with the baryta. The hypophosphite of barium remains in solution, and may be separated from the insoluble phosphate by filtering. Sulphuric acid, added to the solution, precipitates the barium, and leaves the free hypophosphorous acid in solution. By evaporation it may be reduced to the consistency of syrup.

It forms salts called **hypophosphites**, which are all soluble in water.

PHOSPHORUS TRIOXIDE (P_2O_3) is the anhydride of Phosphorous Acid (H_3PO_3), and is obtained by burning phosphorus in a limited supply of air, when a white powder is formed. It

unites with water, forming phosphorous acid. The acid is also obtained by adding water to the trichloride of phosphorus (PCl_3), when mutual decomposition takes place ($\text{PCl}_3 + 3\text{H}_2\text{O} = \text{H}_3\text{PO}_3 + 3\text{HCl}$). The hydrochloric acid is expelled by heating the liquid, and the phosphorous acid is obtained as crystals. This acid exposed to the air absorbs oxygen, and becomes phosphoric acid. When heated it gives off phosphoretted hydrogen gas, and is converted into phosphoric acid.

Its salts are called **phosphites**, but they are of little importance.

PHOSPHORUS PENTOXIDE (P_2O_5), the anhydride of Phosphoric Acid (H_3PO_4) is obtained by burning phosphorus in oxygen, or dry atmospheric air, when a snow white powder is formed, which rapidly absorbs moisture, and runs into a liquid. This white powder has a very great affinity for water, and when thrown into water, unites with it with a hissing noise. Phosphorous acid may also be formed by the action of water on phosphorus pentachloride, thus: $\text{PCl}_5 + 4\text{H}_2\text{O} = \text{H}_3\text{PO}_4 + 5\text{HCl}$.

PHOSPHORIC ACID is a powerful acid; it has a strong, sour taste, reddens litmus paper, and forms an important class of salts called **phosphates**. It is not poisonous, and it is used in medicine under the name of glacial phosphoric acid.

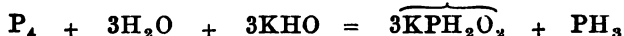
Ordinary phosphoric acid (H_3PO_4), or **orthophosphoric acid**, as it is sometimes called, contains three atoms of replaceable hydrogen, and is capable of uniting with three equivalents of a monatomic base. When this acid is heated to redness it parts with one equivalent of water, and becomes converted into **metaphosphoric acid** (HPO_3), which has only one atom of replaceable hydrogen. When strongly heated for some time, ordinary phosphoric acid becomes converted into **pyrophosphoric acid** ($\text{H}_4\text{P}_2\text{O}_7$), according to the following equation: $2\text{H}_3\text{PO}_4 = \text{H}_4\text{P}_2\text{O}_7 + \text{H}_2\text{O}$, the four hydrogen atoms of which are replaceable by a base.

PHOSPHORETTED HYDROGEN.

Symbol PH_3 *Combining weight* 34.

Phosphoretted hydrogen, as a gas, may be obtained by heating phosphorous acid,— $4\text{H}_3\text{PO}_3 = 3\text{H}_3\text{PO}_4 + \text{PH}_3$. It is generally prepared by heating phosphorus and sticks of caustic potash in a little water, which produce phosphoretted hydrogen and hypophosphite of potassium.

Hypophosphite of
Potassium.



Great care is required in this process. The retort should be small, and have a long tube, which should be placed under water from the first, as the gas formed in this way has the property of taking fire spontaneously in atmospheric air. It is also advisable to displace the air in the retort previous to preparing the gas by sending through it a current of coal gas. Each bubble of the gas, as it escapes into the air, bursts into flame, and changes into a beautiful white wreath of phosphoric anhydride, expanding and revolving as it rises.

The inflammability of the gas prepared in this way is due to the presence of a minute quantity of the vapour of the liquid phosphide of hydrogen (P_2H_4). The gas obtained by heating hydrated phosphorus acid is not inflammable.

Phosphoretted hydrogen has a strong smell of garlic; it is highly poisonous, and slightly soluble in water. When sent through solutions of certain metallic salts, it precipitates phosphides of the metals. It has been suggested that this and similar compounds formed from the decomposition of animal and vegetable matters in marshy places, &c., give rise to the phenomena called by the names—"Will-o'-the-Wisp" and "Jack-o'-Lantern."

There are three compounds of phosphorus and hydrogen, viz.—the gas described above; liquid phosphide of hydrogen (P_2H_4); and solid phosphide of hydrogen (P_4H_2).

Phosphorus combines with chlorine, forming phosphorus trichloride (PCl_3), a colourless liquid, and phosphorus pentachloride (PCl_5), a yellowish crystalline solid. Each of these compounds is decomposed by water, as described above.

ARSENIC is closely allied to phosphorus in its chemical characters, and forms one of a natural group of elements, comprising nitrogen, phosphorus, and arsenic. In this work arsenic is considered along with the metals.

QUESTIONS ON PHOSPHORUS AND ITS COMPOUNDS.

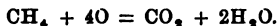
1. How much Anhydrous Phosphoric Acid is formed by the combustion of one gramme of Phosphorus ($P=31$)? How much pure air is needed for the production of this Phosphoric Acid? (L.)
2. What combinations does Phosphorus form (1) with Hydrogen, and (2) with Chlorine? How is each of these substances obtained? (O.)
3. What are the characteristic properties of Phosphorus? How is it obtained? (O.)
4. Give the composition of Phosphorous Acid, and describe the method of preparing it. State also the composition of the Phosphites. *
5. Give the formula of Phosphoretted Hydrogen, its percentage composition by weight and by volume, and its specific gravity. *
6. Describe the preparation of red Phosphorus, and give its properties. (C.)
7. Describe the preparation of gaseous Hydrogen Phosphide (phosphoretted hydrogen). (O.)
8. Give the formula, and a short description of the acids of Phosphorus. (O.)
9. How is Phosphorus obtained from Bone-ash? What weight of Phosphorus is present in 100 grammes of pure Phosphate of Calcium? (O.)
10. Describe the experiments you would perform in order to decide whether the oxide formed by the combustion of phosphorus in oxygen is an acid or a base? (L.)
11. What is the composition of the following substances, viz:—Air, Water, Silica, Marble, Ammonia, Marsh-gas, and Blacklead, or Plumbago? Give the names of the substances represented by the following formulæ:— CO , NO , SO_2 , N_2O_3 , PH_3 , and NH_4Cl (L.)

QUANTITATIVE CHEMICAL PROBLEMS.

I. Correction of gases for temperature and pressure. See pages 13 and 14.

II. Rule for finding the volume of a substance produced in certain chemical decompositions, &c. See page 24.

Ex.—How many cubic feet of oxygen are requisite to consume completely 4 cubic feet of marsh gas, and how many cubic feet of carbon dioxide and water will be formed?

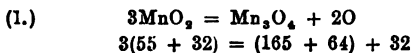


By the law given on page 24, the above equation expresses the fact that 2 volumes of marsh gas combine with 4 vols. of oxygen, to form 2 vols. of carbon dioxide and 4 vols. of steam.

Therefore, 4 cubic feet of CH_4 require, for complete combustion, 8 cubic feet of O; and there are formed 4 cubic feet of CO_2 , and 8 cubic feet of steam.

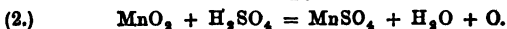
III. To find the weight of a substance employed in a given chemical reaction.

Ex.—How many parts of oxygen may be obtained from 100 grms. of peroxide of manganese—(1) by the action of heat alone, (2) by the action of sulphuric acid? [$\text{Mn} = 55$; $\text{O} = 16$.]



\therefore 261 parts by weight of MnO_2 , yield 32 parts by weight of oxygen, and 100 grms. MnO_2 yield $\frac{100}{261} \times 32 = 12.26$ grms. O.

If the quantity of oxygen by volume is required, then since 16 grms. O occupy 11.2 litres, 12.26 grms. occupy $\frac{11.2 \times 12.26}{16} = 8.5$ litres.



87 parts by weight of MnO_2 , yield 16 parts by weight of O, or 100 grms. of MnO_2 yield $\frac{100 \times 16}{87} = 18.3$ grms. O.

IV. To find the percentage composition of a compound body, or the quantity of a given element contained in a certain quantity of a compound.

Ex. 1.—Find the percentage composition of Nitric Acid (HNO_3).

$$\text{Molecular weight of HNO}_3 = (1 \times 1) + (14 \times 1) + (16 \times 3) = 63.$$

Since 63 parts of Nitric Acid contain 1 of Hydrogen, 14 of Nitrogen, 48 of Oxygen.

$$\begin{array}{rcl} \therefore 1 \text{ part of} & \text{H} & \frac{1}{63} \times 100 = 1.58 \text{ per cent. of H} \\ \therefore 100 \text{ parts of} & \text{N} & \frac{14}{63} \times 100 = 22.22 \\ & \text{O} & \frac{48}{63} \times 100 = 76.19 \end{array}$$

Ex. 2.—How many grammes of Oxygen are contained in a kilogramme of Potassium Chlorate (KClO_3)?

$$\begin{array}{c} \text{K} \quad \text{Cl} \quad \text{O}_3 \\ \text{Molecular weight of Salt} = 39 + 35.5 + 48 = 122.5. \end{array}$$

122.5 grammes of KClO_3 contain 48 grammes of Oxygen.

$$\begin{array}{rcl} \therefore 1 \text{ gramme} & \text{O} & \frac{48}{122.5} \\ \therefore 1 \text{ kilogramme,} & \text{O} & \frac{48}{122.5} \times 1000 = 391.83 \text{ grms.} \\ \text{or 1000 grammes} & \} & \end{array}$$

V. To find the symbol or formula of a body, its percentage composition being given :—

Divide the given percentage number of each element by its atomic weight ; divide each quotient by the lowest, and prefix to the result the symbol of the element.

Ex.—Find the formula of a body, of which the percentage composition is Calcium 38.72 ; Phosphorus 20 ; Oxygen 41.28.

$$\begin{array}{rcl} \text{Relative number of atoms of Ca} & = & \frac{38.72}{40} = .968 \\ \text{P} & = & \frac{20}{31} = .645 \\ \text{O} & = & \frac{41.28}{16} = 2.58 \end{array}$$

Reducing these numbers to their simplest relation, by dividing each by the lowest, we have the number of atoms of Ca = $\frac{.968}{.645} = 1.5$ nearly.

$$\begin{array}{rcl} \text{P} & = & \frac{.645}{.645} = 1 \\ \text{O} & = & \frac{2.58}{.645} = 4 \end{array}$$

This would give the formula of the body as follows :— $\text{Ca}_{1.5}\text{PO}_4$; but as by the atomic theory we can have no frac-

tional parts of atoms entering into combination, we must multiply the numbers obtained as above by the smallest number (in this case, 2), which will bring them all to whole numbers, and we get the formula $\text{Ca}_3\text{P}_2\text{O}_8$, or $\text{Ca}_3(\text{PO}_4)_2$, which is the formula for tricalcium phosphate.

EXERCISES.

1. How many litres of Carbon Dioxide (Carbonic Anhydride) will be formed by the complete combustion of 12 grms. of Carbon, and how many litres of Oxygen will be required? *

2. Calculate the percentage composition of each of the following compounds:—Cupric Sulphate ($\text{CuSO}_4 + 5\text{H}_2\text{O}$); Alum ($\text{KAl}_2\text{SO}_4 + 12\text{H}_2\text{O}$), Nitric Acid (HNO_3), and Potassium Ferrocyanide ($\text{K}_4\text{FeC}_6\text{N}_6$). *

3. What weight of Chlorine is there in one litre of Hydrochloric Acid Gas at 0° and 760 mm? (L.)

4. If 2 volumes of Marsh Gas weighing 16 grms. be burnt in excess of air, how many volumes of carbon dioxide (carbonic acid gas) are produced? Also, what is the weight of the other compound formed at the same time? (L.)

5. How many lbs. of Nitric Acid can be obtained from 150 lbs. of Nitre (potassium nitrate), and how many grms. of Sulphuric Acid will be needed? *

6. What volume of Hydrogen, measured in cubic centimetres at the normal temperature and pressure, is evolved when 39 milligrammes of potassium and 40 milligrammes of calcium are thrown into water? (L.)

7. If you heat common salt with hydric sulphate (H_2SO_4), hydric chloride (HCl) is given off. How many cubic centimetres of hydric chloride at 0° C. and 760 mm. pressure could you liberate from 1000 grms. of sodic chloride? (L.)

8. The action of sodium on water is represented by the equation $\text{Na}_2 + 2\text{H}_2\text{O} = 2\text{NaHO} + \text{H}_2$. How much water is required to take part in this reaction in order to obtain 22.38 litres of hydrogen at 0° C. and 760 m.m. pressure? (L.)

9. Supposing the air to contain 23 per cent. of oxygen by weight, how many grms. of it must be supplied in order to burn completely (a) 10 grms. of carbon, (b) 10 grms. of sulphur, (c) 10 grms. of phosphorus? ($\text{C}=12 : \text{S}=32 : \text{P}=31$). (L.)

10. A mineral has been found to contain in 100 parts, SiHcom 20.28, Magnesium 34.16, Oxygen 45.56. What is its formula? *

11. A compound of Iron and Oxygen contains 72.4 per cent. of iron, and 27.6 per cent. of oxygen. What is its formula? *

12. The analysis of a mineral gives Silica = 55, Lime = 25.6, Magnesia = 18.3. Find its formula: $\text{SiO}_2 = 60$, $\text{CaO} = 56$, $\text{MgO} = 40$. (C.)

13. A cubic inch of Nitrous Oxide weighs .475 grains. What quantity of Nitrate of Ammonia is required to produce 5 cubic feet of this gas? (O.)

14. How much Hydrogen could you obtain by the solution of 100 grms. of Zinc in hydrochloric acid? How much water could be obtained by the combustion of this hydrogen? [At. weight of zinc = 65.2.]

15. What space will 5.324 grms. of Ammonia Gas occupy at 15° C. and 745 m.m. pressure?

16. How much Carbon will be required to convert 10 litres of Carbon Dioxide into Carbon Monoxide, and what will be the volume of the gas so obtained?

17. Calculate the density of Nitrous Oxide (N_2O) at 0° C. and 760 m.m. pressure.

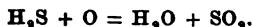
At. wt. of O = 16.

At. wt. of N = 14. (L.)

18. Calculate the percentage composition of Saltpetre from the formula KNO_3 . [K = 39; N = 14; O = 16.]

19. A gas is composed of 92.3 parts of Carbon and 7.7 parts of Hydrogen; how many atoms of hydrogen are united in this gas to 2 atoms of Carbon? (L.)

20. The combustion of Hydric Sulphide in air is represented by the equation



How much Sulphurous Acid and how much Water can be formed from 100 grms. of Hydric Sulphide? [H = 1; S = 32; O = 16.] (L.)

THE METALLIC ELEMENTS.

The metallic elements are much more numerous than the non-metallic, and they are distinguished from the latter by several characteristics, which are only shown, however, when metals are in mass. These distinctions, which are for the most part physical and not chemical, are—(1) a characteristic *lustre*, or power of reflecting light, giving rise to the appearance known as the *metallic lustre*; (2) a greater power of conducting heat and electricity than the non-metals; (3) a greater opacity than the non-metals; and (4), as a rule, a much higher specific gravity than the non-metals. It may also be noted that in the majority of cases the oxides of the non-metallic elements are acid in character, whereas those of the metals are basic. There are some metallic oxides, however, which have an acid character.

All the metals are solid at ordinary temperatures, excepting mercury, which is liquid. They differ from each other in physical and chemical properties, and in uses. The metals employed in the arts in the metallic state are—*Aluminium, Antimony, Bismuth, Copper, Gold, Iron, Lead, Magnesium, Mercury, Nickel, Platinum, Silver, Tin, Zinc*. The properties which render these metals so valuable are the following:—their *lustre, density, tenacity, malleability, and ductility*.

Physical properties. One striking property of the metals is the *lustre* which they acquire by polishing, pressure, &c.

In colour there is considerable uniformity. In most cases it is more or less white, with shades of blue, red, and yellow.

One metal, copper, is red; gold, calcium, barium, and strontium are yellow.

Their specific gravity differs widely, passing from potassium and sodium, which are lighter than water, to platinum, which is twenty-one times heavier. The specific gravities or densities of the more important metals are as follows:—

Platinum	21·5	Iron	7·8
Gold	19·3	Tin.....	7·3
Mercury	13·6	Zinc	7·1
Lead	11·3	Aluminium ...	2·6
Silver.....	10·6	Sodium	0·97
Copper	8·9	Potassium	0·86

The various metals differ greatly in the temperature at which they become liquid or fuse. Mercury is liquid at all temperatures above $39\cdot4^{\circ}\text{C.}$; tin melts at about 228°C. , that is, below a red heat; gold melts at 1100°C. ; cast iron at a temperature of about 1500°C. ; while platinum is infusible with the greatest heat of a blast-furnace, and is fused only by the extreme temperature of the oxyhydrogen flame, or that associated with the electric light.

The property of malleability, or the power of being hammered out into thin sheets, is possessed by certain of the metals. The most malleable is gold, which may be beaten out into leaves having a thickness only of $\frac{1}{250000}$ of an inch. Silver, copper, and tin are malleable in a less degree.

Ductility, or the power of being drawn out into wire, is a property of the malleable metals, and in this there is much difference. Gold, silver, and platinum are all very ductile metals. Iron and copper are also ductile, but to a less extent.

Tenacity, or the power of resisting a force tending to pull the particles asunder, is a property possessed in various degrees by the different metals. Iron, in the form of steel, is by far the most tenacious metal; and among those in common use lead has the least tenacity.

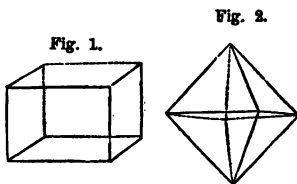
Volatility, or the power of being converted into vapour, is possessed by many of the metals. Mercury boils at 662° F. (350° C.); arsenic passes into vapour before becoming liquid; potassium, sodium, zinc, magnesium, and cadmium rise in vapour at a red heat.

Metals also differ considerably as regards **hardness**; potassium and sodium may easily be cut with a knife; gold and silver in the pure state are also comparatively soft; iron and chromium are very hard metals.

Crystallization. Almost all substances capable of existing in the solid state, under favourable circumstances, assume a definite geometrical form, or are said to **crystallize**. Bodies which do not crystallize are called **amorphous**. The most beautiful examples of crystallization are seen in the natural minerals. As a rule, every substance possesses a definite form in which it always crystallizes, and by which it may be known.

The most common forms are the cube (1), and the regular octahedron (2),—the latter is an eight-sided solid, of which the faces are equilateral triangles. Common salt occurs crystallized in the form of cubes, so does the common mineral known as iron pyrites (disulphide of iron). Alum crystallizes in octahedral forms. The metal bismuth may be obtained in beautiful crystals, which are apparently cubes, though strictly they are rhombohedra.

Certain substances exhibiting a similar chemical constitution, and crystallizing in the same forms, are called **isomorphous**. The various compounds known as “alums” are examples of isomorphous bodies.



CHEMICAL PROPERTIES OF METALS.

Metals combine one with another to form alloys, and with the non-metallic elements to form oxides, sulphides, chlorides, &c.

In the latter case the metallic characters are often lost, whilst in the former case they are preserved.

ALLOYS. As only few of the metals possess qualities which render them suitable to be employed alone, other metals are combined with them, and the compound is called an alloy. Thus gold and silver are each hardened by an addition of a little copper. The hard and useful substance **brass** is an alloy of copper and zinc. **Bronze** is an alloy of copper and tin. **Printing type** is an alloy of lead and antimony.

When the metals combine with mercury, the resulting body is called an **amalgam**.

METALLIC OXIDES. Most metals readily combine with oxygen, and form oxides. The majority of metals burn in air, or pure oxygen; the precious metals, as silver, gold, and platinum, are not combustible, but combine with oxygen indirectly.

The oxides, as pointed out under oxygen, are divided into three classes, **basic, neutral, and acid**. As has been already stated, the metallic oxides are usually basic. Many oxides are found native; others are obtained by combustion in air or oxygen, or by the decomposition of a salt.

METALLIC SULPHIDES. All the metals combine with sulphur to form sulphides, which are frequently found in nature as ores. Many metals unite with sulphur when simply heated with that element. Other sulphides are formed by heating the sulphate of the metal with charcoal, so as to remove its oxygen; others are formed by the action of sulphuretted hydrogen upon some soluble compound of the metal.

The sulphides of the alkalies and alkaline earths are soluble in water; the remainder are insoluble, but some are soluble in acids and alkalies.

METALLIC CHLORIDES. Many metals combine directly with chlorine especially when in a finely divided state. Some

chlorides are found native, as common salt (chloride of sodium). A very convenient way of obtaining a chloride is to add hydrochloric acid to a metallic oxide.

Metals combine with bromine, iodine, fluorine, &c., and these compounds much resemble the corresponding chlorides.

Metals also unite with nitrogen, phosphorus, boron, carbon, silicon, and hydrogen, but these compounds are not important.

DISTRIBUTION AND CLASSIFICATION OF METALS.

Only a few of the metals are found free in nature. Most of them are found in combination with oxygen, sulphur, or other metals. The compounds from which man extracts metals are called *ores*. These ores are usually found in the form of veins, that is, filling cracks in the rocks composing the crust of the earth.

The metals may be conveniently grouped in four classes, viz. :—

I. METALS OF THE ALKALIES. (1) Potassium ; (2) Sodium ; (3) Lithium ; (4) Cæsium ; (5) Rubidium. [(6) Ammonium.]

These are distinguished by forming with oxygen basic oxides, which are very soluble in water, and possess alkaline properties, viz. :—(1) a peculiar bitter taste ; (2) the power of changing vegetable reds into blues or greens, and vegetable yellows into browns ; (3) the power of neutralizing or destroying the characteristic properties of acids, when added to them in sufficient quantity.

II. METALS OF THE ALKALINE EARTHS. (1) Barium ; (2) Strontium ; (3) Calcium.

Their oxides are earthy powders, having alkaline properties, but are less soluble than those of Class I.

III. METALS OF THE EARTHS. (1) Aluminium ; (2) Glucinum ; (3) Yttrium ; (4) Erbium ; (5) Cerium ; (6) Lanthanum ; (7) Didymium.

Excepting aluminium, all these are very rare. Their oxides are earthy powders, insoluble in water.

IV. METALS PROPER, or the metals with which we are most familiar, thirty-four in number.

This class may be further divided into—common metals, as iron, lead, copper, &c., which readily combine with oxygen, and tarnish in air ; and noble metals, as gold, which have only a feeble attraction for oxygen, and do not tarnish in air.

The metal magnesium is sometimes classified with the alkaline earths ; but in several particulars it stands in close relation to zinc and cadmium, and for this reason it is reckoned as one of the metals proper.

I. METALS OF THE ALKALIES.

(1) Potassium ; (2) Sodium ; (3) Lithium ; (4) Cæsium ; (5) Rubidium. [(6) Ammonium.]

(1) POTASSIUM (Kalium).

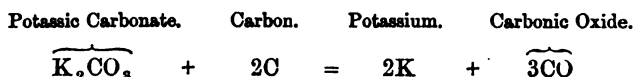
Symbol K. Combining weight 39. Sp. gr. 0.865.

This metal was discovered by Sir H. Davy, in 1807, when he succeeded in decomposing fused potash (hydrated potassium oxide, KHO) by means of a powerful electric current.

Potassium is a bright white metal. At ordinary temperatures it can be cut with a knife, but at 0° C. it is brittle and crystalline. It melts at 62.5° C., and can be distilled at about 420° C. It has great affinity for oxygen, so that if it is exposed to the air, it immediately becomes covered with a film of its oxide. On this account the metal must be preserved in some liquid, as naphtha, which does not contain oxygen. Its

affinity for this gas is best seen by throwing a small piece into water, which it immediately decomposes. One atom of the metal liberates one-half the hydrogen of a molecule of water, and forms the hydrate of potassium (caustic potash), KHO , which dissolves in the water. The heat developed by the intensity of this reaction is sufficient to ignite the liberated hydrogen, which burns with a beautiful violet coloured flame, owing to the presence of vapour of potassium.

Preparation. Potassium compounds occur in nature abundantly, especially in the igneous rocks (granite, trap, &c.), in the form of silicates. Plants extract these compounds from the soil formed from the crumbling of the rocks; and from the ashes of burnt plants, carbonate of potassium (or pot-ashes) is obtained. This carbonate is heated with charcoal, which removes the oxygen, producing carbonic oxide, and liberates the metal potassium. Thus—



Sea water is another source of potassium compounds. The nitrate and chloride of potassium are found in large quantities in various localities.

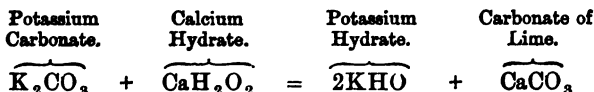
OXIDES OF POTASSIUM. Of these there are three—

1. Potash, or the Suboxide K_2O
2. Dioxide of Potassium K_2O_2
3. Tetroxide of Potassium K_2O_4

The last two of these are of little importance, but the first, combined with water, is an important substance.

POTASH is obtained pure by exposing thin slices of potassium to dry air. It is a white fusible and volatile solid, having great affinity for water. When once united with water, the latter cannot be removed by heat.

POTASSIUM HYDRATE, or **Caustic Potash** (KHO), is obtained by boiling a solution of carbonate of potassium and slaked lime (calcium hydrate): the lime combines with the carbonic acid, forming carbonate of lime, which falls as a heavy powder, and potassium hydrate remains in solution.



The solution is evaporated to dryness, then the residue is melted and cast into moulds. It is sold in the form of white sticks, under the name of caustic potash. This substance is soluble in half its weight of water, and possesses strong caustic properties, destroying both animal and vegetable substances.

Potassium Carbonate, **Potassic Carbonate**, or **Carbonate of Potassium** (K_2CO_3), known in commerce as **pot-ashes**, or **pearl-ashes**, is brought in large quantities from America and Russia. The ashes of burnt wood are boiled with water, and the solution evaporated to dryness, when the residue forms the crude pot-ashes; and these, when further purified, are called pearl-ashes. The pure carbonate may be obtained from these. A pure carbonate is also obtained by heating cream of tartar to redness, dissolving the mass in water, and evaporating the filtered liquid to dryness.

Carbonate of potassium is very soluble in water, and has strong alkaline properties. It is largely consumed in the arts, —in the manufacture of soap, glass, in preparing caustic potash, and other compounds of potassium.

A bicarbonate of potassium (KHCO_3) is formed by passing a current of carbonic acid through a solution of carbonate of potassium.

Potassium Nitrate, or **Nitrate of Potassium** (KNO_3), well known as **nitre**, or **saltpetre**, is one of the most important salts of potassium. It is brought from the East Indies, where it is found in the soil. The soil is boiled with water, and the filtered

solution evaporated, when the nitre crystallizes. It is also largely prepared artificially.

Saltpetre readily parts with its oxygen on the application of heat, and for this reason it is used in the manufacture of gunpowder.

Gunpowder is a **mechanical mixture** of nitre, sulphur, and charcoal, in the proportion of 75 parts of nitre, 15 of charcoal, and 10 of sulphur. When gunpowder is fired, the oxygen of the nitre unites with the charcoal, forming carbonic acid and carbonic oxide, the nitrogen is set free, and the sulphur unites with the potassium. The great explosive power of gunpowder is due to the sudden formation of these gases, which, on account of the heat produced, have a volume 1500 times greater than the volume of gunpowder employed.

Potassium forms several compounds with sulphur, of which the best known is the mono-sulphide (K_2S).

The other important salts of potassium are **chloride of potassium** (KCl),—a salt extracted from kelp, and largely used in the manufacture of alum: **iodide of potassium** (KI),—a salt much used in medicine: **bromide of potassium** (KBr): **chlorate of potassium** ($KClO_3$): and **sulphate of potassium** (K_2SO_4).

Almost all the potassium salts are soluble in water.

TESTS FOR POTASSIUM.

1. A solution of tartaric acid gives, after briskly stirring with a glass rod, a white precipitate of the acid tartrate of potassium (cream of tartar).

2. A solution of perchloride of platinum ($PtCl_4$) produces a yellow crystalline precipitate of the double chloride of platinum and potassium ($2KCl + PtCl_4$).

Blow-pipe test. The dry salt introduced upon platinum wire into the **reducing flame** of the blowpipe, communicates to it a violet tint. The colour of the flame may be seen by throwing saltpetre into an open fire.

(2) SODIUM (Natrium).

Symbol Na. *Combining weight* 23. *Sp. gr.* 0.972.

Sodium is found in nature abundantly, but always in a state of combination. Its most important compound is common salt, the chloride of sodium, which is found in Cheshire, Galicia, &c., and also in sea water. Compounds of sodium are found in several minerals, and in land and sea plants.

This metal was discovered by Sir H. Davy, about 1807, in the same manner as potassium, viz., by the decomposition of the fused hydrate (caustic soda), by means of a powerful electric current. It has a silver-white colour, is soft at ordinary temperatures, and melts at 95.6°C . It burns with a bright yellow flame. When thrown upon water, it decomposes a portion of it, like potassium, liberating the hydrogen, which, however, does not take fire unless the water be hot; sodium hydrate (NaHO) is formed, and dissolves in the water.

Preparation. Sodium is prepared from its carbonate in the same way as potassium. On account of its affinity for water, it must be kept in naphtha. It is much cheaper than potassium.

OXIDES OF SODIUM. There are two well-known oxides, viz :—

Oxide of Sodium, or Soda... Na_2O

Dioxide ,, Na_2O_2 (not used in the arts).

Soda is obtained anhydrous by burning sodium in dry air, a white powder being formed. It takes up moisture readily, and forms the hydrate (NaHO).

The hydrate of sodium, or **caustic soda**, is obtained by boiling lime with the carbonate of sodium, and then evaporating the filtered solution to dryness. It is a white solid substance, highly soluble in water, caustic, alkaline, and closely resembles potash in its properties. It is largely used in soap-making.

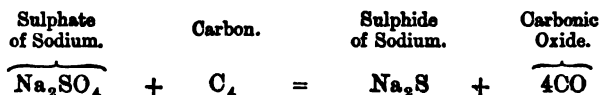
Chloride of Sodium, common salt (NaCl), is the substance from which the sodium compounds are prepared. It is found in sea water, &c., as mentioned above. It is soluble in about

three parts of cold water. It is largely used in the arts, and for domestic purposes.

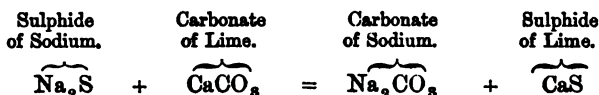
SULPHATE OF SODIUM, or **Glauber's salt** ($\text{Na}_2\text{SO}_4 + 10\text{H}_2\text{O}$), is obtained by the action of oil of vitriol upon common salt. It has a bitter taste, and is used in medicine.

CARBONATE OF SODIUM, ($\text{Na}_2\text{CO}_3 + 10\text{H}_2\text{O}$), is a very important compound, largely used in glass-making, soap-making, bleaching, medicine, and for domestic purposes. It was formerly obtained from the ashes of sea plants, but it is now manufactured from the sulphate of sodium.

This sulphate, obtained from common salt by the action of oil of vitriol, is heated in a furnace with chalk and powdered coal. Two stages of decomposition take place. In the first stage, the sulphate parts with all its oxygen to the coal, and becomes the sulphide: thus,—



In the second stage, the sulphide parts with its sulphur to the lime of the chalk, whilst the carbonic acid of the chalk unites with the sodium of the sulphide, and becomes carbonate of sodium: thus—



The soluble carbonate is separated from the insoluble sulphide by treating the mass with water,—then evaporating the carbonate solution to dryness. The residue is roasted, and is known as the **soda-ash** of commerce. If the soda-ash be dissolved, and the solution allowed to stand, large crystals of the hydrated carbonate of sodium ($\text{Na}_2\text{CO}_3 + 10\text{H}_2\text{O}$) will be formed.

The carbonate is very soluble in water, and has strong alkaline properties.

Bicarbonate of Sodium (NaHCO_3) is obtained by exposing the carbonate to the action of carbonic acid gas. A white powder is formed, having a feeble alkaline action, which is used in medicine and for effervescing drinks.

4. **PHOSPHATES OF SODIUM.** There are several of these, but the one best known, and from which most of the phosphates are formed, is the **hydro-disodic phosphate**, better known as rhombic phosphate of sodium ($\text{Na}_2\text{HPO}_4 + 12\text{HO}$). It is a bitter, purgative salt, used in medicine. It is also useful as a test.

5. **NITRATE OF SODIUM** (NaNO_3), known as **Cubic Nitre**, and **Chili Saltpetre**, is found in large beds in Peru. It is used in the manufacture of nitric and sulphuric acids, and also for manure.

Other compounds of sodium are **borax**—(see boron); the **sulphide** of sodium (Na_2S); the **bromide** (NaBr); the **iodide** (NaI); and various **silicates**.

TESTS FOR SODIUM.

Nearly all the sodium salts are soluble in water, therefore there is no good direct test for this metal. In analysis, if the metal be proved to be one of the alkalies, the presence of sodium is shown by the absence of potassium and ammonium.

Before the blow-pipe, salts of sodium give a bright yellow colour to the flame.

(3) LITHIUM.

Symbol Li. *Combining weight* 7. *Sp. gr.* 0.59.

Lithium is of recent discovery, and is named from "lithos" ($\lambda\iota\theta\omicron\varsigma$)—a stone, because it was found at first only in minerals. It is a white metal, like sodium, and is the lightest solid

known. It is widely distributed in nature ; in small quantities it is found in micas and felspars, in tobacco, and in mineral waters. The metal is obtained by decomposing its chloride by electricity.

Its oxide is called lithia (Li_2O). This metal forms several compounds, as the chloride, the sulphate, the phosphate, and carbonate, all of which communicate a purplish red colour to the flame of a Bunsen lamp.

(4) CÆSIUM.

Symbol Cs. Combining weight 133.

(5) RUBIDIUM.

Symbol Rb. Combining weight 85. Sp. gr. 1.52.

These metals were discovered about 1860, by means of spectrum analysis ; they are widely distributed, and found in small quantities in various rocks, mineral springs, and in the ashes of several plants, viz., beetroot, tobacco, coffee, and grapes. They closely resemble potassium in their properties.

(6) AMMONIUM.

Symbol NH_4 or Am. Combining weight 18.

The compound radical, or group of atoms NH_4 , forms a series of salts, which very closely resemble those of potassium and sodium, and it is usual, therefore, to class them under the head of alkalis. The radical (which is called ammonium) forms an amalgam with mercury, when sodium or potassium amalgam is poured into a solution of ammonium chloride. The radical itself is said to have been prepared in the free state, and to consist of a dark blue liquid body, having a metallic lustre.

Ammonium hydrate (NH_4Ho or AmHo) is formed when ammonia gas is dissolved in water. It has alkaline properties resembling those of the corresponding compounds, KHO , and NaHO . It is a valuable re-agent in the laboratory.

COMPOUNDS OF AMMONIUM.

The most important are as follows :—

1. **SULPHIDE OF AMMONIUM**, $(\text{NH}_4)_2\text{S}$, or Am_2S , is produced by the direct combination of sulphuretted hydrogen and ammonia gas (NH_3), at a low temperature; at ordinary temperatures the sulphide loses ammonia and becomes converted into the hydro-sulphide $(\text{NH}_4)\text{HS}$. A solution of this latter compound in water, prepared by passing sulphuretted hydrogen gas through a solution of ammonia, is a valuable re-agent in the laboratory. When freshly prepared it is colourless, but it soon turns yellow.

2. **CHLORIDE OF AMMONIUM**, $(\text{NH}_4)\text{Cl}$, or AmCl , is the most important salt of ammonium. It is commonly called **muriate of ammonia**, and **sal-ammoniac**. It is formed directly by the union of hydrochloric acid and ammonia gas; but it is obtained chiefly from the ammoniacal liquor of gas works by the addition of hydrochloric acid. It is usually sold in the shape of large hollow cakes, but it may also be obtained in crystals. It is very tough and fibrous, and very soluble in water. It readily volatilizes when heated. Its sp. gr. is 1.578.

3. **SULPHATE OF AMMONIUM**, $(\text{NH}_4)_2\text{SO}_4$, is obtained from ammoniacal liquor in the same way as the chloride, by adding sulphuric acid instead of hydrochloric acid. It is used as a fertilizer.

4. **NITRATE OF AMMONIUM**, (NH_4NO_3) , is obtained by neutralizing nitric acid with ammonia. It is important as the source of nitrous oxide gas.

5. **CARBONATES OF AMMONIUM**. There are several compounds of carbonic acid and ammonium. The most important of these is the **sesqui-carbonate**, the common carbonate, or **sal volatile** and **smelling salts** of the shops. The composition of this compound is somewhat complex and uncertain; it is

represented by the formula $(\text{NH}_4)_4\text{C}_3\text{O}_8$, and is regarded by some chemists as a compound of ammonium carbonate, and carbon dioxide.

This carbonate is prepared by heating together powdered chalk and chloride of ammonium. It has a strong pungent smell, like that of ammonia, an acrid taste, and a strongly alkaline reaction. It is much used in the arts and in medicine:—all the salts of ammonium may be obtained from it by adding the respective acids.

TESTS FOR AMMONIUM.

1. Ammonium is easily detected by the odour of ammonia, which is given off from all its salts on the addition of quick-lime or potash.

2. A glass rod dipped in hydrochloric acid produces dense white fumes when brought into the vapour of ammonia.

There is no blow-pipe test, as most of the salts of ammonium, when heated, are volatilized.

THE ALKALIES are distinguished from each other, thus:—

Tartaric Acid gives, with **Potassium**, a white precipitate.

” ” **Sodium**, no precipitate.

” ” **Ammonium**, no precipitate.

Ammonium and sodium are distinguished by the smell of ammonia on the addition of caustic potash.

The alkalies are distinguished from the alkaline earths by the following test:—**Phosphate of sodium** and ammonia together produce in the latter **white precipitates** of the phosphates; in the former, **no precipitate**.

II. METALS OF THE ALKALINE EARTHS.

There are three of these, viz:—**Barium**, **Strontium**, and **Calcium**. As **Magnesium** resembles these metals in some of its properties, we shall consider it under this head.

These metals differ from the alkalis in forming insoluble carbonates and phosphates.

(I) BARIUM.

Symbol Ba. *Combining weight* 137. *Sp. gr.* 4.5.

This metal was obtained by Sir H. Davy, in 1808, by decomposing certain of its compounds by a powerful electric current. It has a yellowish-white colour, decomposes water at ordinary temperatures, and tarnishes rapidly when exposed to the air. It occurs abundantly in nature in the form of sulphate and carbonate.

The name Barium is derived from the Greek word βαρύς (heavy), on account of the great weight of its compounds.

OXIDES OF BARIUM. Of these there are two:—1. The protoxide **baryta** (BaO). 2. The dioxide (BaO_2). The last is of little importance, except for the preparation of hydrogen peroxide (H_2O_2).

Baryta (BaO). This oxide is obtained anhydrous by heating the nitrate of barium in a crucible till the acid is expelled. It is a grey porous mass, slakes like lime when moistened with water, and forms the hydrate of baryta (BaOH_2O , or BaH_2O_2). If hot water be saturated with this hydrate, and allowed to cool, a crystalline hydrate is formed, having the composition ($\text{BaH}_2\text{O}_2 + 8\text{H}_2\text{O}$). This hydrate is soluble in 20 parts of cold water:—the solution is alkaline and poisonous: and on exposure to the air absorbs carbonic acid, and turns milky.

SALTS OF BARIUM. These are chiefly interesting as tests for the mineral acids, especially for sulphuric acid. The two most important are chloride of barium (BaCl_2), and the nitrate of barium $\text{Ba}(\text{NO}_3)_2$. These are obtained by dissolving the carbonate or sulphide of barium in hydrochloric acid and nitric acid respectively. Chloride of barium is soluble in water, and this solution is the usual test for sulphuric acid. With sul-

phuric acid and soluble sulphates it forms a white precipitate of barium sulphate, BaSO_4 , which is insoluble in water and dilute acids.

The other important salts are the sulphate of barium (BaSO_4), which is found native as **heavy spar**, a most insoluble salt, prepared artificially, and used as a white paint; the sulphide (BaS), formed from the sulphate by heating **heavy spar** with coal, and dissolving in water; the carbonate (BaCO_3), which occurs native as the mineral **witherite**.

The soluble salts of barium are poisonous. The tests for barium compounds will be considered with those for the compounds of the metals of the other alkaline earths.

(2) STRONTIUM.

Symbol Sr. *Combining weight* 87.5. *Sp. gr.* 2.54.

This metal resembles Barium in its properties. Its oxide was discovered about the year 1787, and was so called from Strontian, in Argyleshire, where strontium carbonate (**Strontianite**) occurs native in the lead mines. It is less abundant than barium. The metal is malleable, of a pale yellow colour, and is prepared by the action of electricity upon the fused chloride. It burns in air with a crimson flame, and forms the oxide strontia (SrO).

Strontia may be obtained by heating the nitrate. It forms with water a hydrate ($\text{SrO} + 9\text{H}_2\text{O}$), which is soluble in water, and the solution has similar properties to baryta water

SALTS OF STRONTIUM. The two most important are the chloride of strontium (SrCl_2), and the nitrate of strontium $\text{Sr}(\text{NO}_3)_2$, which are soluble in water. They are used in the manufacture of fireworks to give a crimson colour to flame.

The salts found native are the carbonate (SrCO_3), and the sulphate (SrSO_4), known in mineralogy by the names of **strontianite** and **celestine** respectively. These native salts are insoluble in water.

(3) CALCIUM.

Symbol Ca. *Combining weight* 40. *Sp. gr.* 1.58.

The compounds of this metal are among the most abundant and important constituents of the earth's crust. It is found in fluor-spar, in combination with fluorine; in limestone and chalk, in combination with carbonic acid; in gypsum, combined with sulphuric acid. The metal was discovered by Davy, in 1808; it is obtained from the fused chloride by means of electricity. It has a light yellowish colour, is very malleable, and combines with oxygen on exposure to the air, forming its only oxide, lime (CaO). The metal, when heated, burns in the air with a brilliant light, and decomposes water at the ordinary temperature. It derives its name **Calcium** from the Latin, **Calx** (lime).

COMPOUNDS OF CALCIUM.

OXIDE OF CALCIUM, or **Lime** (CaO), is obtained by heating limestone (the carbonate of calcium) in kilns; the carbonic acid is driven off, and lime remains. Pure lime, or **quicklime**, is a white, porous, infusible substance; it readily combines with water, giving off great heat, and forming the hydrate, or **slaked lime** ($\text{CaO}, \text{H}_2\text{O}$, or CaH_2O_2). The hydrate is soluble in about 700 parts of cold water, and in about 1300 of hot water, so that a saturated solution of lime-water loses nearly half its lime on boiling. Lime-water has all the properties of solutions of potash and soda in a smaller degree. It is used in the laboratory as a test for carbonic acid.

Lime is chiefly used in making mortar and cements. **Mortar** is a mixture of slaked lime and sand; the lime, as it dries, absorbs carbonic acid from the air, and hardens on account of the formation of carbonate of calcium. Sand is used in the composition of mortar, to prevent the lime crumbling and falling to pieces as it dries; but, besides this, a gradual combination takes place between the lime and the silica of the sand,

leading to the formation of silicate of lime, which contributes much to the hardness of mortar.

Ordinary mortar cannot be used under water. Mortars for this purpose, called hydraulic mortars, are obtained by using lime containing clay and silica.

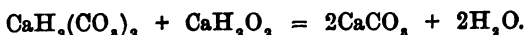
Lime is also used as a manure. On rich vegetable soils it is useful in destroying the excess of vegetable matter; on heavy clay soils it liberates the alkalies, especially potash, which is especially valuable in connection with the growth of plants.

CARBONATE OF CALCIUM, or as it is more commonly called **carbonate of lime** (CaCO_3), is found abundantly in nature, under the forms of chalk, limestone, marble, coral. It is the chief constituent of the shells of molluscs, egg shells, and coral reefs, and is also a constituent of bone. It is found in the crystalline form as Calc spar, or Iceland spar. It is soluble in pure water to the extent of only two grains in one gallon, but it is freely taken up by water containing carbonic acid, owing to the formation of a soluble bicarbonate, or calcium dihydrogen carbonate, $\text{CaH}_2(\text{CO}_3)_2$. Water thus charged with carbonate of lime is said to be **hard**. As the excess of carbonic acid escapes by exposure to the air, evaporation of the water, or is expelled by heat, carbonate of lime is deposited. In this way those masses of deposits called **stalactites** and **stalagmites**, found in limestone caves, are formed.

Most spring waters contain carbonate of lime, held in solution by excess of carbonic acid. During boiling this acid is expelled, and the well-known crust of carbonate of lime is formed inside the boiler. This evil is prevented by adding sal-ammoniac to the water, which produces the soluble chloride of calcium, and volatile carbonate of ammonium.

Hard waters, which owe their hardness to the presence of the soluble bicarbonate of lime, may be softened by boiling, or by the addition of lime water to neutralize the excess of carbonic acid. This latter method is employed on the large scale in

some parts of England, and is known by the name of Clark's process. The re-action is expressed by the following equation :—



The hardness of water which can thus be removed is termed **temporary hardness**, to distinguish it from permanent hardness of water, which is due to the presence of sulphate of calcium.

SULPHATE OF CALCIUM, or **sulphate of lime**, (CaSO_4), is found **anhydrous** in a crystalline form in the Tyrol and Austria; but it is most abundant as a hydrate in combination with two atoms of water. In this state it is called **gypsum** and **alabaster**, and the clear crystalline varieties are termed by the mineralogist **selenite** ($\text{CaSO}_4 + 2\text{H}_2\text{O}$). When **gypsum** is carefully heated, it parts with its water, and forms **plaster of Paris**. This, when moistened, takes up again two atoms of water, and returns to a solid mass, and on that account is used for moulds and casts.

Sulphate of calcium is soluble in water to the extent of 150 grs. per gallon. The solution is useful for a test to distinguish the alkaline earths from each other. This salt is a common impurity of spring water, and gives to it the property of permanent hardness.

CHLORIDE OF CALCIUM (CaCl_2) is obtained by dissolving limestone or marble in hydrochloric acid. By evaporation of the solution, the chloride is obtained in crystals associated with six atoms of water: four of these atoms are expelled by gentle heat, and the salt, in this form, has such an affinity for water, that it is largely employed for the purpose of drying gases.

Chloride of calcium must not be confounded with the substance commonly known as **chloride of lime** (CaOCl_2). The latter is also known as **bleaching powder**, and is obtained by exposing the hydrate of lime, very slightly moist, to chlorine gas.

PHOSPHATES OF CALCIUM. There are several of these compounds, but the most interesting is **bone phosphate**, or **tricalcium phosphate**, $\text{Ca}_3(\text{PO}_4)_2$, which is the chief constituent of the bones of animals. This phosphate may be obtained by digesting calcined bones in nitric acid, and precipitating the filtered solution by ammonia.

FLUORIDE OF CALCIUM, or **FLUOR-SPAR** (CaF_2), is mentioned under fluorine. Other compounds are the **phosphide** (CaP_2), and the **sulphide** (CaS).

(4) MAGNESIUM.

Symbol Mg. *Combining weight* 24. *Sp. gr.* 1.74.

This metal is found abundantly in nature in combination with carbonate of lime, as **magnesian limestone**. It occurs in sea water as the chloride, and in many springs as the sulphate. It also occurs in the form of a silicate in many important minerals (steatite, talc, meerschaum, serpentine, &c)

The metal is prepared by heating the **chloride** along with **sodium**, when chloride of sodium is formed, and metallic magnesium.

In properties, this metal closely resembles zinc and cadmium. It is of a silver-white colour, malleable and ductile. It does not tarnish in dry air; but it slowly oxidizes in moist air. It burns in air with a dazzling white light, and, on account of this property, it is used in photography as a substitute for sunlight.

COMPOUNDS OF MAGNESIUM.

Magnesia (MgO) is the only known oxide of this metal, and may be formed by burning magnesium in air. It is usually obtained by heating the carbonate, or nitrate, and in this form it is largely used in medicine, under the name of calcined magnesia. It is a white, tasteless, soft, almost insoluble powder, with feeble alkaline properties. It forms a hydrate with water.

SULPHATE OF MAGNESIUM ($\text{MgSO}_4 + 7\text{H}_2\text{O}$) is the most important compound. It may be obtained from sea water by evaporating the **bittern** along with sulphuric acid, or by the action of sulphuric acid upon magnesian limestone. This salt is a common ingredient in mineral waters, and because of its occurrence in the springs about Epsom, it is known as **Epsom salts**. It is soluble in three times its weight of water; by its solubility it is distinguished from the sulphates of the three preceding metals.

CARBONATE OF MAGNESIUM (MgCO_3) is found native as a white, hard mineral, called **magnesite**. It is not easily prepared artificially. The **magnesia-alba** used in medicine is a mixture of the carbonate and hydrate, and is prepared by adding a hot solution of sodium carbonate to a boiling solution of magnesium sulphate.

CHLORIDE OF MAGNESIUM (MgCl_2) is interesting as being the compound from which the metal is obtained.

PHOSPHATES OF MAGNESIUM. The most important phosphate is that known as the **phosphate of magnesium and ammonium** ($\text{MgNH}_4\text{PO}_4 + 6\text{H}_2\text{O}$). It is formed, as a crystalline white precipitate, by adding any soluble salt of phosphoric acid, and ammonia, to a solution of sulphate of magnesium. The formation of this salt furnishes a delicate test for the presence of magnesium.

TESTS FOR THE ALKALINE EARTHS.

Sulphate of Lime produces, with salt of Barium, a white precip.		ediately.
„	„	Strontium, a white precip. after a time.
„	„	Calcium or Magnesium, no precip.
Oxalate of Ammonium	„	Calcium, a white precip.
„	„	Magnesium, no precip.

With the blow-pipe, Barium gives to a flame a yellowish-green colour.

„ Strontium „ crimson colour.
 „ Calcium „ orange-red „

All the compounds of the alkaline earths, when heated strongly in the outer flame of the blow-pipe, become highly luminous. Magnesium compounds, moistened with nitrate of cobalt sheeted on charcoal, take a pale-rose colour.

GENERAL QUESTIONS ON THE ALKALIES AND ALKALINE EARTHS.

1. Name the Alkaline Earths, and give their principal chemical characters. (C.)
2. How is Ammonia-gas usually prepared in the laboratory? Express by symbols the reactions that occur. What weight of Sal-ammoniac must be taken to produce 100 grammes of Ammonia-gas? (C.)
3. How is Sodium Carbonate manufactured for the purposes of commerce? Explain the process, using symbols. How many tons of pure Sodium Carbonate can be, according to theory, obtained from 100 tons of common salt? (C.)
4. Compare the compounds of Ammonia, Potash, and Soda, pointing out the several characters in which they resemble one another, and those in which they differ. (C.)
5. What weight of Sodium Carbonate must be dissolved in a litre of water, that 2000 parts by weight of the solution may contain Na_2CO_3 , or one molecule of the salt? (O.)
6. Who discovered Sodium, and by what means? Describe that substance. (C.)
7. Describe and explain the preparation of a solution of Caustic Soda. (C.)
8. Explain the production of the "fur" in tea-kettles in which hard water is boiled. (C.)
9. Explain the difference between Limestone, Quicklime, and Slaked Lime. (O.)
10. How is Hyposulphite of Soda prepared? For what purpose is it chiefly used? (C.)
11. What is the composition of Ammonia—(1) by weight, (2) by volume. (C.)
12. How is Potassium obtained, and what are its properties? (O.)

13. How is Lime obtained? What are its chemical properties? (O.)
14. What weights of Sulphuric, Nitric, and Hydrochloric Acids, respectively, are required to neutralise 212 parts by weight of Carbonate of Sodium? What is the weight of the solid product in each case? (O.)
15. What relative weights of Carbonate of Calcium, Nitrate of Ammonium, and Chlorate of Potassium must be taken to produce equal volumes of Carbonic Acid, Nitrous Oxide, and Oxygen? ($\text{Ca} = 20$, $\text{Cl} = 35.5$, $\text{K} = 39.1$). (O.)
16. Give a process by which Barium Salts may be distinguished from Salts of Calcium. *
17. You boil common Well-water for a few minutes, and observe the separation of a white precipitate. How do you explain the formation of the precipitate? How would you test your explanation by experiment? (L.)
18. Ammonia gas is led into a solution of hydrochloric acid. What is produced? Give an equation. The resulting liquid is evaporated to dryness in a porcelain crucible, which is afterwards heated to redness. What will remain in the crucible? *
19. State the composition of Gunpowder, and explain the reactions which take place in its combustion? *
20. The spring water of our chalk districts is much harder than the river water of the same localities; why is this the case? At Canterbury the water is rendered soft by the addition of a certain quantity of lime water; explain the process. (L.)

III. METALS OF THE EARTHS.

ALUMINIUM.

Symbol Al. *Combining weight* 27.5. *Sp. gr.* 2.6.

Aluminium occurs abundantly in nature in the form of oxide (alumina), in clays, and many other varieties of rocks. It also occurs combined with fluorine and sodium in **cryolite** ($3\text{NaF}, \text{Al}_2\text{F}_6$), a mineral occurring abundantly in Greenland, and it is now prepared in large quantities from that mineral. The metal was discovered in 1828, but has only been used commercially within the last few years.

The metal may be obtained from its chloride (Al_2Cl_6) in the same manner as magnesium. It is of a bluish-white colour, malleable, and ductile, and resembles zinc both in colour and

hardness. It is very sonorous. On account of its lightness and lustre it is used in the manufacture of ornaments and trinkets. An alloy, composed of 90 parts of copper and 10 of aluminium, resembles gold in appearance, and is largely manufactured under the names of aluminium bronze, Abyssinian gold, &c. When heated in a current of oxygen, aluminium burns with a brilliant bluish-white light, producing its oxide—alumina. The metal is not acted on by exposure to the air at ordinary temperatures, and this fact, together with its white colour, the high polish which it is capable of receiving, and its lightness, has caused it to be much used in various philosophical instruments. It is readily soluble in hydrochloric acid.

Alumina (Al_2O_3) is the only known oxide of this metal. It occurs abundantly in combination with silicon and oxygen in felspar, also in all varieties of clay, and in every kind of soil. It is found native, crystalline and almost pure, as the sapphire, ruby, corundum, and emery. All these crystallized forms of alumina are remarkable for their extreme hardness, which is surpassed only by that of the diamond.

Alumina may be prepared by adding ammonium hydrate to a solution of alum, when the hydrate ($\text{Al}_2\text{O}_3 + 3\text{H}_2\text{O}$) is thrown down as a white, gelatinous substance. The hydrate, on being heated, parts with its water, and forms a white soft powder of pure alumina. The hydrate is readily soluble in acids, and in the caustic alkalies; but when anhydrous, alumina is scarcely acted on by acids. In either state, it is insoluble in water; it has no taste, and does not change the colour of litmus. It acts as a weak base. Alumina and its compounds are of great use to the dyer and calico printer. Hydrated alumina has the property of forming insoluble compounds, called lakes, with certain colouring matters. Fast colours are given to cloth by soaking it first with a preparation of alumina, and then dipping it in the colouring matter. The alumina adheres to the fibre of the cloth as well as to the colouring matter, so that the latter cannot be washed out.

COMPOUNDS OF ALUMINIUM.

The **CHLORIDE OF ALUMINIUM** (Al_2Cl_6) is of interest as being one of the sources of the metal. It is obtained as a yellow, anhydrous, volatile substance, by heating a mixture of alumina and charcoal in dry chlorine gas.

SULPHATE OF ALUMINIUM ($\text{Al}_2(\text{SO}_4)_3 + 18\text{H}_2\text{O}$) is obtained by dissolving alumina in sulphuric acid. It is made on a large scale under the name of alum cake, by the action of sulphuric acid upon clay. It is much used by the dyer and calico printer. It has a great tendency to unite with other sulphates, giving rise to a series of compounds called alums.

COMMON POTASH ALUM is a salt composed of the sulphates of aluminium and potassium, which crystallizes in regularly shaped octahedral crystals, containing a large quantity of water of crystallization. It has the composition represented by the formula $\text{K}_2\text{Al}_2(\text{SO}_4)_4 + 24\text{H}_2\text{O}$. It may be prepared by mixing solutions of the two sulphates, and allowing the compound salt to crystallize. Alum is manufactured on the large scale from decomposing shale or clay containing iron pyrites (FeS_2), which is found abundantly near Whitby and Glasgow, and called alum schist, or alum ore. On roasting the shale, oxygen is absorbed from the air, and sulphuric acid is produced, which, uniting with the alumina of the clay, forms sulphate of aluminium. Sulphate of iron, which is also formed at the same time, remains in solution with the sulphate of aluminium, but it is decomposed on the addition of chloride of potassium, sulphate of potassium being formed, and alum crystallizes out.

Potash alum has an astringent taste; it is soluble in 18 parts of cold water, and in its own weight of boiling water, and the solution has a strongly acid reaction, and dissolves iron and zinc. Alum is employed in medicine, and for dyeing.

AMMONIA ALUM is an exactly similar compound, containing sulphate of ammonium instead of sulphate of potassium; it is

prepared in large quantities, and used in the arts. Other alums known are soda alum, cesium alum, rubidium alum, silver alum, &c. All the alums are very soluble in water, have an acid reaction, and an astringent taste.

SILICATES OF ALUMINIUM. These are numerous and important compounds. The various clays, formed by the crumbling away of rocks, consist chiefly of these compounds. Clay is really the result of the action of air and water upon rocks containing felspar. Clay is the chief constituent of earthenware, porcelain, or china.

Common earthenware is made of a mixture of clay and silica, dried in the air, and baked in a kiln. It is firm, solid, porous, brittle, and gives out a dull sound when struck. To make earthenware fit to contain liquids, it is glazed by coating it with a fusible glass, and then heating it in a kiln till the surface is covered with a glassy varnish. For the commoner kinds of earthenware, common salt is used to produce the glaze.

Porcelain or china is made of the finest kind of clay, derived directly from the decomposition of felspar, and called **kaolin**, or china clay. There are various kinds of porcelain, according to the recipes of the makers.

Porcelain differs from earthenware in breaking with a glassy fracture, in possessing a ringing sound when struck, and in being semi-transparent. Like earthenware, it must be glazed to make it fit to hold liquids.

TESTS FOR ALUMINIUM COMPOUNDS.

1. Sulphide of ammonium gives, in a solution of aluminium compounds, a **white** precipitate of the **hydrate of alumina**, with evolution of sulphuretted hydrogen. Only one other metal—zinc, gives a white precipitate with this test.

2. Caustic potash produces a **white** precipitate of hydrated alumina, soluble in excess. In this solution **chloride of ammonium**, or **ammonia**, gives a white precipitate of hydrated

alumina; while in a similar solution, sulphuretted hydrogen, or ammonium sulphide, gives no precipitate. This test distinguishes aluminium from zinc.

With the blow-pipe a salt of aluminium, moistened with nitrate of cobalt, and heated on charcoal, acquires a pale blue colour.

The metals of the alkalies and alkaline earths are distinguished from the other metals by giving no precipitate with sulphide of ammonium.

Glucinum, Yttrium, Erbium, Cerium, Lanthanum, Gallium, and Didymium are very rare metals, prepared in the same way as aluminium, and forming compounds like it. The carbonates are insoluble in water, but the sulphates are soluble.

IV. METALS PROPER.

Div. I.—Metals not precipitated in acid solutions by Sulphuretted Hydrogen:—Zinc, Iron, Manganese, Cobalt, Nickel, Chromium.

(1) ZINC.

Symbol Zn. Combining weight 65. Sp. gr. 6·8 to 7·1.

The ores of zinc are the carbonate, called **calamine**; the sulphide, called **zinc blende**, or **black-jack**; the red oxide, called **red zinc ore**; and the hydrated silicate, known as **electric calamine**.

The metal is extracted by roasting the powdered ores. During this process, the sulphur and the carbon burn away as sulphurous acid and carbonic acid respectively, and the zinc becomes oxidized. The roasted ore is then mixed with fine coal or charcoal, heated in retorts of peculiar shape, and metallic zinc is distilled over, and condenses.

Properties. Zinc is a bluish-white metal of crystalline structure. It is brittle at ordinary temperatures, but when

heated to a temperature from 100°C. to 150°C. it becomes ductile and malleable. Heated still higher it becomes brittle again. It melts at about 412°C. ; at red heat it rises in vapour, and if this be exposed to the air it burns with a bluish-white flame, forming oxide of zinc. It tarnishes in a moist atmosphere, becoming coated with a thin film of the oxide. Unlike the oxide or rust which is formed on iron exposed to the air, the thickness of this film of zinc oxide does not go on increasing, but the coating first formed protects the metal under it from further action. It dissolves in dilute acids, giving off hydrogen.

Zinc, on account of its durability, cheapness, and lightness, is used instead of lead in roofing. It is also used in the construction of galvanic batteries. Iron covered with zinc is preserved from oxidation, and is said to be **galvanized**. Zinc is known in commerce as "spelter."

Alloys of Zinc and Copper. A number of alloys of these two metals are known under the names of **brass**, **Muntz's metal**, **pinch-beck**, **Dutch metal**, while **German silver** is an alloy formed of copper, zinc and nickel. **Brass** varies somewhat in composition; in general it is composed of 2 parts of copper to 1 of zinc. **Dutch metal** is an alloy consisting of 11 parts copper to 2 parts zinc, and is so malleable as to be capable of being beaten out into thin leaves, which are used in many cases as a substitute for gold leaf. **German silver** or **nickel silver** varies considerably in composition, that used for spoons and forks is said to be composed of 2 parts copper, 1 part nickel, and 1 part zinc.

Monoxide of Zinc (ZnO). This is the only known oxide. It is obtained by burning the metal in air, or by precipitating a soluble salt of zinc by means of an alkali, and heating the precipitate. It is a white, insoluble powder. When heated it becomes yellow, but on cooling it returns to its original white colour. The oxide is soluble in acids.

Sulphide of Zinc (ZnS), when precipitated from an alkaline

solution by sulphuretted hydrogen, is of a white colour. **Blende**, one of the most abundant zinc ores, is generally of a black colour, from its admixture with sulphide of iron, and is hence called **black-jack**.

Chloride of Zinc (ZnCl_2), a white, soluble, very deliquescent substance, is obtained by heating zinc in chlorine gas, or by dissolving the metal in hydrochloric acid. A solution of this salt under the name of **Burnett's disinfecting fluid** is largely used as an antiseptic and disinfectant, and as a preservative of wood and vegetable fibre against decay.

Sulphate of Zinc ($\text{ZnSO}_4 + 7\text{H}_2\text{O}$), a soluble salt, the **white vitriol** of commerce, obtained by dissolving zinc in sulphuric acid. It is used in medicine, and in calico printing.

Carbonate of Zinc (ZnCO_3), an insoluble compound, found native, both massive and crystallized, as calamine. It is usually of a greyish or yellowish colour.

TESTS FOR ZINC.

1. Sulphuretted hydrogen gives no precipitate in acid solutions.

2. Sulphide of ammonium gives a **white** precipitate of hydrated sulphide of zinc, readily soluble in hydrochloric acid.

3. Caustic potash gives a **white** precipitate of hydrated oxide of zinc, soluble in excess. Tests 2 and 3 have the same reaction with aluminium compounds. In the potash solution, sulphuretted hydrogen gives a **white precipitate** of the hydrated sulphide of zinc, whereas aluminium gives no precip.

Blow-pipe tests. 1. A salt of zinc, mixed with carbonate of sodium on charcoal, and exposed to the reducing flame, yields a **yellow incrustation**, which becomes white on cooling. 2. Moistened with nitrate of cobalt, and heated on charcoal in the outer flame, a salt of zinc acquires a **green** colour.

(2) IRON (Ferrum).

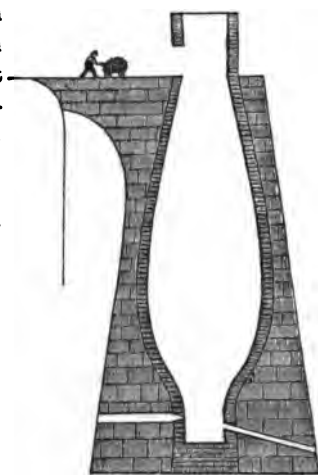
Symbol Fe. Combining weight 56. Sp. gr. 7.8.

Iron is more widely diffused than any other metal, but it rarely occurs native. Its chief ores are **magnetic iron ore**, the black oxide (Fe_3O_4), containing in its pure state over 72 per cent. of iron, and occurring abundantly in Norway, Sweden, and the United States; **specular iron ore**, or red hæmatite (Fe_2O_3), containing when pure 70 per cent. of iron, and found in Cumberland, United States, Sweden, &c.; **brown iron ore**, or **brown hæmatite** ($2\text{Fe}_2\text{O}_3 + 3\text{H}_2\text{O}$), a hydrated form of red hæmatite, containing about 60 per cent. of iron, occurring in Northamptonshire and the district of the Forest of Dean; **spathic iron ore**, or **carbonate of iron** (FeCO_3), found in certain parts of England, and on the Continent, in Westphalia, Styria, Carinthia, &c. But the chief source of iron in this country is **clay ironstone**, an impure carbonate of iron containing about 30 to 33 per cent. of metallic iron, mingled with clay, oxide of manganese, lime, and magnesia. This clay ironstone occurs in beds or in detached nodules, in connection with the beds of coal in N. and S. Wales, Yorkshire, Derbyshire, Warwickshire, &c.

Metallic iron is found in meteoric stones.

Extraction of Iron. The process of extracting iron from those ores in which it is combined only with oxygen is very simple, and consists in heating the ore in contact with carbon, by which means the oxygen is removed from the iron. Brown hæmatite and pure spathic iron ore may be reduced in like manner, if first roasted to expel water and carbonic acid. But in the case of ores containing foreign substances (clay, &c.) the process is more complex, since it is necessary to heat the ore to a very high temperature, in contact with lime, &c., in order to remove those foreign substances. Clay ironstone is first roasted, to drive away water, carbonic acid, and volatile sub-

stances. The roasted ore is then placed in the **blast furnace**, which is a large conical structure about 50 feet high. The ore is introduced near the top of the furnace, along with coal and limestone. At the bottom there are blast pipes, through which heated air is forced, to keep up a high temperature within the furnace. During the process of **smelting**, the limestone separates the clay (silicate of aluminium) from the oxide of iron, forming a glassy compound of silicate of calcium and aluminium, called **slag**: the



oxide of iron parts with its oxygen to the coal, and falls to the bottom of the furnace, as liquid iron. The liquid slag and iron are drawn off at intervals,—the latter running into channels of sand, and in this form it is known as **pig**, or **cast iron**.

Cast iron is not pure iron. It contains about five per cent. of carbon, and small quantities of silicon, sulphur, manganese, phosphorus, &c. ; hence there are several varieties of cast iron. Pig iron is highly crystalline and brittle.

Wrought or **Malleable Iron** is formed from cast iron by removing its impurities. This is done by melting it in a furnace of the shape of a baker's oven, in contact with oxide of iron, and freely exposed to the oxidising action of the air. During the process of **puddling**, as it is called, the carbon, &c., is removed, and the melted iron thickens, and is rolled into lumps. These lumps are taken out and hammered to expel the liquid slag, and any remaining liquid cast iron, and, at the same time, coherence is given to the metal. The iron is again melted and rolled into bars, to make the fibres parallel, and thus greater toughness is acquired.

Wrought iron is not quite pure: it contains about $\frac{1}{300}$ of its weight of carbon, and about $\frac{1}{100}$ of silicon. It is fibrous in texture, and very tough; hence it is used for all purposes where a great strain is required.

Wrought iron, when exposed to continued vibration and concussion, loses its fibrous texture and becomes crystalline, and consequently brittle. This change accounts for the breaking of railway-carriage axles, &c., in certain cases.

Steel is made from wrought iron by adding a certain proportion of carbon to the latter. Bars of wrought iron and charcoal powder are placed in iron boxes, and exposed to a red heat for 40 or 50 hours. At the end of that time, the iron will have taken up from 1.3 to 1.5 per cent. of carbon. It is now more fusible than before; it has lost its fibrous texture, and has become fine grained.

The most important character of steel is its power of assuming great hardness and elasticity when heated to redness and suddenly cooled. By varying the heat, and the rapidity of cooling, different degrees of hardness and brittleness are given to steel. This process is called **tempering**.

BESSEMER'S process of manufacturing steel is much quicker and cheaper than that described above. Atmospheric air is sent through fused cast iron, which burns away nearly all the carbon and silica. To the wrought iron thus made, pure molten cast iron is added in sufficient quantity to convert the whole into steel. In this way six tons of cast iron may be converted into steel in twenty minutes.

Properties. Pure iron has a white colour, and a lustre almost equal to silver. It is softer than bar iron, but very tough. In bar iron the colour is bluish-white or grey. It is the most tenacious of all metals. It is strongly magnetic up to a red heat. Iron does not oxidise or tarnish in dry air; heated to redness it becomes covered with a scaly coating of black oxide (Fe_3O_4), and at a white heat it burns brilliantly,

producing the same oxide. Polished iron does not tarnish in pure water, but if carbonic acid be present, and air finds access, the iron rusts, forming the hydrated sesquioxide ($2\text{Fe}_2\text{O}_3 + 3\text{H}_2\text{O}$). The layer of oxide (rust) first formed does not protect the subjacent metal, as in the case of zinc, but the process of oxidation (rusting) goes on continually until the whole mass of iron is transformed into oxide. Iron at a red heat decomposes water, liberating hydrogen, and forming the black oxide. It is soluble in dilute sulphuric and hydrochloric acids, with evolution of hydrogen.

COMPOUNDS OF IRON.

Iron forms two series of compounds, known respectively as **ferrous** and **ferric** compounds. In the ferrous compounds iron appears to be divalent, combining with one atom of oxygen, two atoms of chlorine, &c. In the ferric compounds two atoms of iron occur in combination together, forming a group which acts as a hexad, combining with six atoms of chlorine, &c.

OXIDES OF IRON. Of these there are three, viz :—

- (1) **Protoxide, or Ferrous Oxide** (FeO), the base of the **green** or **ferrous** salts.
- (2) **Sesquioxide, or Ferric Oxide** (Fe_2O_3), the base of the **red** or **ferric** salts.
- (3) **Black or Magnetic Oxide** (Fe_3O_4), which does not form any definite salts.

1. **FEROUS OXIDE** has a great affinity for oxygen, passing into a higher oxide, hence the difficulty of obtaining it pure. Its hydrate is obtained by adding caustic potash to a solution of a ferrous salt, when a **white** precipitate is thrown down, which changes to a **greenish-brown** colour if oxygen be present. This oxide colours glass green ; and all its salts, when pure, are of a pale green colour.

The most important ferrous salts are—

• **Ferrous Sulphide, or Sulphuret of Iron** (FeS), prepared by

bringing a bar of iron, heated to whiteness, in contact with roll sulphur. It is largely used in the preparation of sulphuretted hydrogen.

Disulphide of Iron (FeS_2), found abundantly in nature as iron pyrites; it is used in the production of sulphuric acid.

Ferrous Chloride, or Protochloride of Iron (FeCl_2), a soluble salt, formed by passing dry hydrochloric acid gas over metallic iron, or by dissolving iron in hydrochloric acid.

Ferrous Sulphate, or Protosulphate of Iron ($\text{FeSO}_4 + 7\text{H}_2\text{O}$), a soluble salt, commonly called **green vitriol** and **copperas**. It is prepared by dissolving pure iron, or ferrous sulphide, in oil of vitriol. On evaporation large green crystals are formed, which are used in making black dyes and writing ink. This salt has great affinity for oxygen.

Ferrous Carbonate, or Protocarbonate of Iron (FeCO_3), an insoluble salt, found native abundantly in clay ironstone. It is contained in all ferruginous springs. When the carbonic acid which held it in solution escapes, hydrated ferric oxide is deposited, hence those ochrey deposits around such springs.

2. FERRIC OXIDE, the Sesquioxide, Peroxide, or Red Oxide (Fe_2O_3)—occurs native in great abundance as one of the ores of iron. It is obtained **anhydrous** by heating green vitriol; in this state it is of a red colour, **rouge** being one of its varieties. The **hydrate**, obtained by precipitating it from a solution of one of the ferric salts by means of caustic potash, is of a brown colour.

The salts of this oxide are all of a brown colour. The chief ferric salts are—

Ferric Chloride, or Perchloride of Iron (Fe_2Cl_6), obtained in anhydrous reddish-brown scales by passing dry chlorine gas over red-hot iron. The hydrate is obtained in red crystals by dissolving ferric oxide in hydrochloric acid.

Ferric Sulphate, or Persulphate of Iron ($\text{Fe}_2(\text{SO}_4)_3$), obtained by the action of strong sulphuric acid on ferric hydrate.

3. **BLACK or MAGNETIC OXIDE OF IRON** (Fe_3O_4) is found native as the mineral **loadstone**; it is black in colour, and has a metallic lustre.

This is the oxide formed by heating iron in air or oxygen.

4. **FERRIC ACID.** When iron is fused along with saltpetre, a compound known as potassium ferrate is obtained, in which an oxide of iron plays the part of an acid. This oxide has the formula Fe_2O_6 , but has never been isolated.

TESTS FOR IRON.

Iron forms two classes of salts, **protosalts**, or ferrous salts, and **persalts**, or ferric salts, which are readily distinguished from each other by the colour of their solutions. The **proto-salts** in solution and crystallized have a **pale green** colour; the **per-salts** have a **yellow** or **reddish-brown** colour.

TESTS FOR A FERROUS SALT.

1. Sulphuretted hydrogen in acid solutions gives **no precip.**
2. Sulphide of ammonium gives a **black** precip. of the protosulphide, which dissolves in dilute hydrochloric acid.
3. Caustic potash gives a **whitish-green** or **grey** precip. of the hydrated protoxide, passing quickly into **brown** on exposure to the air.
4. Ferrocyanide of potassium gives a **pale blue** precip., deepening in colour on exposure.
5. Ferricyanide of potassium gives a **deep blue** precip. of the ferricyanide of iron.

TESTS FOR A FERRIC SALT.

1. Sulphuretted hydrogen gives a **white** deposit of sulphur.
2. Sulphide of ammonium gives a **black** precip. of sulphide of iron and sulphur.
3. Caustic potash gives a **brownish-red** precip.
4. Ferrocyanide of potassium gives a **deep blue** precip. of Prussian blue.

5. Ferrieyanide of potassium gives no precip., but the solution turns reddish-brown.

6. Sulphocyanate of potassium produces a blood-red solution.

With the blow-pipe both classes of salts act alike. They give a bottle green colour to borax in the reducing flame, which becomes yellowish in the oxidising flame.

(3) MANGANESE.

Symbol Mn. Combining weight 55. Sp. gr. 8.

The ores of this metal are abundant. The most important ore is the black oxide. The metal is not used in the arts, and its preparation is a mere matter of curiosity. It may be obtained by heating the oxide very strongly with charcoal. It is a metal of a reddish-white colour, brittle, and hard enough to scratch glass. It oxidises on exposure to the air, decomposes water at the ordinary temperature, and should be kept in sealed tubes, or under naphtha.

The compounds of the metal are useful in the laboratory for preparing hydrochloric acid, chlorine, and as a source of oxygen. They are also used in colouring glass, in the preparation of cast steel, and in dyeing.

OXIDES OF MANGANESE. The most important are the

Protoxide, or Manganous Oxide (MnO).

Sesquioxide, or Manganic Oxide (Mn_2O_3).

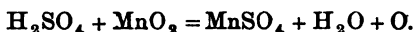
Black, Manganic Dioxide, Binoxide, or Peroxide (MnO_2).

The Protoxide (MnO), obtained by heating the carbonate in a current of hydrogen, is an olive green powder, which absorbs oxygen from the air, and which, on ignition, changes into a higher oxide of a brown colour. Its hydrate ($MnO + H_2O$) is precipitated as a whitish gelatinous body when an alkali is added to a soluble manganese compound. The hydrate rapidly deepens in colour, consequent upon absorbing oxygen. This oxide forms with acids pink-coloured salts.

The Sesquioxide (Mn_2O_3) occurs anhydrous in nature as

braunite. This is the oxide into which the protoxide changes on exposure to air. It is a weaker base than the corresponding oxide of iron, or alumina. The violet colour given to glass, and the colour of the amethyst, are due to this oxide.

The **Black Oxide** (MnO_2) is the most abundant ore of manganese. In a pure state the native oxide is known as **pyrolusite**. When heated, it gives off one-third of its oxygen, leaving a red oxide (Mn_2O_3). If heated with concentrated sulphuric acid, one-half its oxygen is expelled, with the formation of manganous sulphate. Thus,—



It is largely used in the preparation of chlorine and bleaching powder.

Manganic and Permanganic Acids. If an oxide of manganese be fused in the air in contact with potash (or soda) a green compound is obtained, which is soluble in water, and which has the composition K_2MnO_4 , and is called **potassium manganate**. The oxide of manganese (MnO_3), which here plays the part of an acid, has not been isolated. When manganate of potash (or soda) is dissolved in a large quantity of water and allowed to stand, or a few drops of acid added to the solution, hydrated manganese dioxide is precipitated, and the colour of the solution changes from green to purple. From this solution crystals of a salt known as **potassium (or sodium) permanganate**— $\text{K}_2\text{Mn}_2\text{O}_8$ —are obtained on evaporation. The compounds termed permanganates readily part with their oxygen, and are therefore powerful oxidising agents; on this account they are frequently used for disinfecting purposes. A solution of potassium permanganate in water is known as **Condy's Disinfecting Fluid**.

The **Chloride of Manganese** ($\text{MnCl}_2 + 4\text{H}_2\text{O}$) is the residue left in the preparation of chlorine from the **black oxide** and hydrochloric acid. Iron is usually present; but the pure

chloride, when crystallized by evaporation, is of a delicate pink colour.

The **Sulphate of Manganese** ($\text{MnSO}_4 + 5\text{H}_2\text{O}$) is obtained by the action of sulphuric acid on the black oxide. It is a very soluble salt, of a beautiful rose colour, largely used by the dyer to give a permanent brown dye.

TESTS FOR MANGANESE.

1. Sulphuretted hydrogen gives no precipitate.
2. Sulphide of ammonium gives a flesh-coloured precipitate of sulphide of manganese, which becomes brown on exposure to air, and which is very soluble in acids.
3. Caustic potash gives a white precip. of hydrated protoxide, which changes to brown on exposure to air.

Blow-pipe tests. 1. A compound of manganese, melted with borax in a loop of platinum wire, gives in the outer flame a violet-coloured bead, which becomes colourless in the inner flame.

2. Fused on platinum foil with carbonate of soda, a bluish-green mass is produced.

(4) COBALT.

Symbol Co. *Combining weight* 58.6. *Sp. gr.* 8.95.

This metal is commonly found in combination with arsenic and sulphur. It is not easy to obtain it pure, as its ores are very complex. It is of a reddish-grey colour, hard, strongly magnetic, and very tenacious. It is feebly acted upon by hydrochloric and sulphuric acids; and it does not oxidise on exposure to the air. Cobalt in the metallic state is not used in the arts; many of its compounds are remarkable for the beauty and brilliancy of their colour. They are used for pigments, and give to glass a beautiful blue tint.

OXIDES OF COBALT. There are three oxides, the **protoxide** (CoO), the **sesquioxide** (Co_2O_3), and **cobaltic oxide** (Co_3O_4).

The **protoxide** or **monoxide** is a brown powder, prepared by carefully heating the hydrate, precipitated in a solution of one of the cobalt salts on the addition of an alkali. It is soluble in acids, and forms solutions which have a blue colour when concentrated, but become pink on dilution. It is largely used for colouring porcelain. The **sesquioxide** forms no salts; it is a black, insoluble powder. **Cobaltic oxide** is formed when cobaltous oxide or nitrate is ignited in the air.

The most important salts of cobalt are the **chloride** and the **nitrate**. The **chloride** (CoCl_2) is obtained hydrous, as ruby-red crystals, by dissolving the protoxide in hydrochloric acid; when anhydrous, the crystals are blue. A weak solution of this salt is of a pink colour, and is used as a sympathetic ink. The writing, dried in the air, is invisible, but becomes blue when heated. On exposure to moisture the characters again become invisible.

The **Nitrate of Cobalt** ($\text{Co}(\text{NO}_3)_2 + 6\text{H}_2\text{O}$), prepared by dissolving the oxide in nitric acid, is useful in blow-pipe testing. In the outer blow-pipe flame this salt yields, with alumina compounds, a blue colour; with zinc compounds, green; with those of magnesium, a pale pink.

TESTS FOR COBALT.

1. Sulphide of ammonium gives a black precipitate of cobaltous sulphide (CoS), insoluble in excess, and soluble with difficulty in hydrochloric acid.

2. Caustic potash gives a blue precip., which becomes olive-green on exposure, and rose-red when heated.

3. Before the blow-pipe, the borax-bead acquires a deep blue colour when heated either in the outer or inner flame.

(5) NICKEL.

Symbol Ni. *Combining weight* 58.6. *Sp. gr.* 8.8.

This metal is always found associated with cobalt, and is very closely allied to cobalt in its chemical characters. Its

most abundant ore is **kupfernickel** (NiAs , arsenide of nickel). An artificial product called **speise**, consisting of fused nickeliferous ores, and containing nickel combined with arsenic, is also employed as a source of the metal. The process of extraction of the metal from its ores is complicated. It is prepared chiefly for the manufacture of German silver, which is an alloy of nickel, zinc, and copper.

Pure nickel is a brilliant, silver-white, hard, ductile metal, strongly magnetic—a property which it loses if heated beyond 250°C . It becomes oxidized if exposed to air of a high temperature. Sulphuric and hydrochloric acids dissolve it, but slowly.

OXIDES OF NICKEL. There are two oxides, the **protoxide** (NiO), and the **sesquioxide** (Ni_2O_3); only the first forms salts.

The **protoxide** may be obtained anhydrous by heating the carbonate or nitrate of metal in a covered crucible; or as a **hydrate**, by adding caustic potash to a solution of a nickel salt. The oxide has a greenish colour, and the nickel salts it forms have a delicate green colour.

The most important salts of nickel are the **sulphate** ($\text{NiSO}_4 + 7\text{H}_2\text{O}$), and the **chloride** (NiCl_2).

TESTS FOR NICKEL.

1. Sulphuretted hydrogen gives no precipitate.
2. Sulphide of ammonium gives a black precip.
3. Caustic potash gives a pale green precip. (hydrated oxide of nickel), insoluble in excess.
4. Ammonia gives a pale green precip., soluble in excess, forming a blue solution.

Blow-pipe tests. Nickel salts give with borax in the outer flame a reddish-yellow glass; and in the inner flame a dark grey opaque bead.

(6) CHROMIUM.

Symbol Cr. Combining weight 52.5. Sp. gr. 6.8.

Chromium is found in combination with iron as **chrome iron ore** (FeCr_2O_4 , that is, magnetic oxide of iron in which two atoms of iron are replaced by two atoms of chromium) chiefly in North America and Sweden; and also combined with lead, as **chromate of lead**—a very beautiful mineral, from which the metal was first obtained in 1797.

Chromium is not used in the arts, but its compounds are much used by the painter and dyer.

Pure chromium is the most infusible of all the metals, and is scarcely attacked by the strongest acids.

OXIDES OF CHROMIUM. There are four of these, closely corresponding to those of iron, viz., a **protoxide** (CrO), a **sesquioxide** (Cr_2O_3), a **chromo-chromic oxide** ($\text{CrO}, \text{Cr}_2\text{O}_3$, or Cr_3O_4), and **chromic anhydride**, or **chromic acid** (CrO_3). Of these the most important are the **sesquioxide** and **chromic anhydride**.

The **Sesquioxide** or **Chromic Oxide** (Cr_2O_3) is obtained anhydrous, of a fine green colour, by heating to redness the chromate of mercury. It is obtained as a greyish-green hydrate, by boiling alcohol with a solution of bichromate of potassium, acidulated with sulphuric acid. This oxide forms soluble salts of a green or violet colour, sweet to the taste, and poisonous.

The colouring ingredient of greenstone, the emerald, and other minerals, consists of chromic oxide.

Chromic Anhydride—Chromic Acid, (CrO_3) is easily obtained by mixing equal parts of a solution of bichromate of potassium and oil of vitriol, when chromic anhydride separates in the form of crimson-red crystals. It is very soluble in water, and the solution then has an acid reaction, forming with bases salts called **chromates**. Organic matter deprives it of

half its oxygen, and reduces it to the sesquioxide. To this acid the ruby owes its colour.

THE SALTS OF CHROMIUM are numerous: the most important soluble salts are the **chromate** and **bichromate of potassium**.

Chromate of Potassium (K_2CrO_4) is the source of all the preparations of chromium. It is obtained by heating chrome ironstone in a furnace along with nitrate or carbonate of potassium. The product, digested in water, gives a yellow solution, which, on evaporation, deposits yellow crystals.

Bichromate of Potassium ($K_2Cr_2O_7$) is prepared by adding to a solution of the **chromate** a little sulphuric acid. One half of the potassium is removed by the acid, and the other half remains in combination with two atoms of chromic acid; thus the **chromate** is converted into the **bichromate**. This salt crystallizes in beautiful red crystals. It is largely manufactured for the preparation of chrome pigments.

Chromate of Lead ($PbCrO_4$) is an insoluble salt, prepared by precipitating a solution of acetate of lead by bichromate of potassium. It is the **chrome yellow** of the painter. From this a **bichromate** of lead of a splendid scarlet colour may be obtained by boiling the **chromate** with lime water.

Chromate of Silver (Ag_2CrO_4) is an insoluble salt of a ruby-red colour when crystallized. It is prepared by mixing solutions of nitrate of silver and bichromate of potassium.

Chromium forms compounds with chlorine, of which a very interesting one is **chloro-chromic acid** ($CrO_2 + Cl_2$). It is prepared by fusing 10 parts of common salt with 17 parts of chromate of potassium, and then gently heating in a retort with 30 parts of oil of vitriol. It distils over as a deep red vapour, condensing to a dense red liquid, much like bromine. The vapour is more irritating than chlorine; if the liquid be dropped into a strong solution of ammonia or alcohol, it takes fire.

Sulphate of chromium combines with alkaline sulphates to form a series of compounds known as **chrome alums**, which crystallize in precisely the same forms as common potash or ammonia alum.

TESTS FOR CHROMIUM COMPOUNDS.

1. Ammonium sulphide gives a bluish green precip. of hydrated sesquioxide.

2. Caustic potash gives a green precip., (which dissolves with a green colour in excess) in salts of the sesquioxide.

3. The **chromates** in solution give, with a salt of lead, a **yellow** precip.; with nitrate of silver a **red** precip.; with nitrate of mercury an **orange** precip.

Blow-pipe test. Compounds of chromium fused with borax give an **emerald green** bead; fused on the platinum foil with nitrate of potash, they form yellow chromate of potash.

DIV. II. Metals precipitated by Sulphuretted Hydrogen in neutral, acid, or alkaline solutions, viz.:—Copper, Lead, Silver, Mercury, Bismuth, and Cadmium.

(1) COPPER (Cuprum).

Symbol Cu. *Combining weight* 63.5. *Sp. gr.* 8.93.

This useful metal is found native in North America, and in the state of ore in Cornwall, Chili, South Australia, and Spain. The most common ore is copper pyrites, a compound of sulphur, copper, and iron ($\text{Cu}_2\text{S}, \text{Fe}_2\text{S}_3$). Next in abundance to this is the sulphide of copper (Cu_2S), and less abundant are the carbonates of copper or malachite, and the red and black oxides.

Copper is extracted on the large scale from the carbonate or oxide, by reducing the ores with carbon and silica in a furnace. The process of extraction from copper pyrites is very complicated. The chief seat of copper smelting is Swansea.

Properties. Copper is the only red metal. It is rather hard, very tenacious, ductile, and malleable, and is one of the best conductors of heat and electricity. It does not oxidise in

air at ordinary temperatures, but if heated to redness a layer of oxide is formed upon the surface. Copper melts at a temperature of about 1100°C . Pure water does not affect it; but sea water and solutions of chlorides corrode it. Nitric acid dissolves it readily. Oil of vitriol in the cold does not affect it, but if heated along with it, it is decomposed, sulphate of copper being formed, and sulphurous anhydride being given off. Hydrochloric acid, when heated with copper, dissolves it with the formation of cuprous chloride, and the liberation of hydrogen.

It forms very useful alloys. With zinc it forms brass; with tin it forms bronze, bell-metal, gun-metal, and speculum metal.

COMPOUNDS OF COPPER.

Copper is really a dyad element, but it forms two sets of compounds, one in which it assumes its true dyad character, and the other in which it appears to be monovalent. The former compounds are called **cupric**, the latter **cuprous** compounds.

OXIDES OF COPPER. There are two well defined oxides, which are found native, viz., the **red or cuprous oxide** (Cu_2O) and the **black or cupric oxide** (CuO).

Cuprous Oxide, the Red or Suboxide of Copper (Cu_2O) occurs native in ruby-red crystals. It may be prepared by heating together equal parts of copper filings and the black oxide of copper in a covered crucible, or by boiling a solution of sulphate of copper with sugar and caustic potash. This oxide forms colourless salts with acids, which absorb oxygen from the air, and pass into compounds of the cupric oxide. It is chiefly used to stain glass a beautiful ruby colour.

Cupric Oxide, or the Black Oxide of Copper (CuO), is obtained by heating copper to redness in air, or by heating to redness nitrate of copper in an earthen crucible; or by precipitating the hydrate from one of its salts, by means of caustic

potash. This oxide imparts to glass a beautiful green colour. It is soluble in acids, forming salts of a green or blue colour.

Sulphate of Copper, Cupric Sulphate, or Blue Vitriol ($\text{CuSO}_4 + 5\text{H}_2\text{O}$), may be obtained by boiling copper with sulphuric acid, diluted with half its quantity of water. It forms large blue crystals, soluble in water. The water of crystallization is expelled by heating the salt to 392°F ., when it takes the form of a white powder. Heated to bright redness, the salt is decomposed. This sulphate is largely used in dyeing, and in the preparation of pigments.

Nitrate of Copper, Cupric Nitrate ($\text{Cu}_2\text{NO}_3 + 6\text{H}_2\text{O}$), is a corrosive, soluble salt, obtained in the form of deep blue crystals by dissolving copper in nitric acid, and evaporating the solution. Heat decomposes it, expelling the nitric acid, and leaving the black oxide.

Carbonate of Copper has never been obtained pure; it is always associated with the hydrated oxide of copper. In this form it is found as the mineral **chessylite** ($2\text{CuCO}_3 + \text{CuO}$, H_2O), and as the beautiful mineral **malachite** ($\text{CuCO}_3 + \text{CuO}, \text{H}_2\text{O}$), which is much used for ornamental purposes.

Chloride of Copper, Cupric Chloride (CuCl_2), is obtained by the action of chlorine gas upon copper; or by dissolving the black oxide in hydrochloric acid, and evaporating, when hydrated crystals of a green colour are formed ($\text{CuCl}_2 + 2\text{H}_2\text{O}$). It is very soluble in water and alcohol; a solution of the latter burns with a splendid green flame. **Cuprous chloride** (Cu_2Cl_2), is formed by dissolving cuprous oxide in hydrochloric acid. It has the property of dissolving carbonic oxide (CO).

All the salts of copper are very poisonous. Raw white of egg is a good antidote.

TESTS FOR COPPER.

As the cuprous salts are rarely met with, the following tests apply only to the higher or cupric compounds:—

1. **Sulphuretted Hydrogen** gives a black precip. of sulphide of copper (CuS), insoluble in dilute acids.

2. **Caustic Potash** gives a pale blue precip., which becomes black on boiling.

3. **Ferrocyanide of Potassium** gives a red-brown precip., soluble in ammonia.

4. Polished iron wire, or a strip of clean zinc, introduced into a solution of a copper salt, becomes covered with a deposit of metallic copper.

Blow-pipe tests. A salt of copper, heated with carbonate of soda on charcoal in the inner flame, yields metallic copper, which is easily known by its colour.

A borax bead containing copper, after heating in the oxidising flame, is green while hot, blue when cold. In the reducing flame it is red when cold.

(2) LEAD (Plumbum).

Symbol Pb. Combining weight 207. Sp. gr. 11.36.

The chief ore of lead is **galena**—a sulphide of the metal usually containing some silver—which is principally found in England and Spain.

In extracting the lead from galena, a little lime is added to the ore, in order to remove any siliceous matter it may contain. The process of extraction is as follows:—The crushed ore is first roasted in a reverberatory furnace, through which a strong current of air is allowed to pass. During this roasting, a portion of the sulphide is oxidized, and becomes converted into the sulphate, the sulphur of another portion is burnt away as sulphurous anhydride leaving oxide of lead, whilst another portion remains undecomposed. When the roasting is considered complete, the furnace is closed, and the heat suddenly raised. The oxide and sulphate of lead decompose the remaining sulphide, and the whole mass is converted into metallic lead and sulphurous anhydride.

Properties. Lead is a bluish-white metal, very soft, highly malleable, but not very ductile. It melts at 334°C ., and at a high temperature it absorbs oxygen from the air, and gives off white fumes of oxide. It slowly tarnishes on exposure to the air. Pure water has no action upon the metal, but water charged with air corrodes it, and the oxide of lead thus formed dissolves in the water. If there be any chlorides or nitrates present in the water, the corrosion of the metal is greatly increased. Water containing sulphates or carbonates, as most kinds of spring water, does not become contaminated, since sulphuric acid and carbonic acid both form insoluble salts of lead; but if it contain carbonic acid in excess, a soluble carbonate of lead is formed. On account of the use of lead pipes and lead cisterns, and the poisonous nature of lead compounds, the action of water upon the metal deserves special attention.

Nitric acid dissolves lead even in the cold, with the production of nitrate of lead and the liberation of nitric oxide; hydrochloric acid has little or no action on the metal, and cold dilute sulphuric acid has scarcely any action on it; but boiling strong sulphuric acid converts lead into sulphate (insoluble), and sulphurous anhydride is set free.

Lead forms many useful alloys. **Small shot** is an alloy of lead and arsenic; **type metal** is an alloy of lead and antimony; **pewter** and **solder** are alloys of lead and tin.

COMPOUNDS OF LEAD.

OXIDES OF LEAD. There are four of these, viz., a **suboxide** (Pb_2O), which is of no importance; a **protoxide** (PbO); a **dioxide** or **peroxide** (PbO_2); and a **red oxide** (Pb_3O_4).

The **Protoxide of Lead—Plumbic Oxide** (PbO)—is obtained by heating lead in a current of air. If heated below the point of fusion, a yellow powder is formed called **massicot**; but if the metal be fused, it forms a scaly mass of a brick red colour, called **litharge**. Boiled with oil, it forms sticking-plaster. This oxide is slightly soluble in water, but more soluble in solu-

tions of the caustic alkalies. It combines with bases, and forms an important class of salts. It is also used in the manufacture of flint glass, and in the making of glaze for common earthenware. This oxide is very easily reduced by heating it with charcoal.

Peroxide of Lead—Plumbic Dioxide, (PbO_2)—is usually obtained, as a brown powder, by digesting powdered red lead in boiling nitric acid, diluted with four or five times its bulk of water. It is insoluble in water and acids. When heated, it parts with half its oxygen, and becomes a protoxide. Hot hydrochloric acid acts upon it, liberating chlorine, and forming chloride of lead.

Red Oxide—Red lead, or Minium—(Pb_3O_4) is sometimes considered as a combination of the two preceding oxides ($2\text{PbO} + \text{PbO}_2$), and may be obtained by heating massicot to a low redness, whilst a current of air is allowed to pass over it. It is much used in glass making, paper staining, and as a pigment. Nitric acid decomposes it, forming nitrate of lead, leaving a residue of the dioxide (PbO_2).

NITRATE OF LEAD. Plumbic Nitrate (Pb_2NO_3 , or $\text{Pb}(\text{NO}_3)_2$) is easily obtained by dissolving lead, or the protoxide, in diluted nitric acid. Milk-white crystals are formed, which are soluble in about eight parts of cold water. Heated to redness it decomposes, evolving oxygen and red fumes of N_2O_4 .

ACETATE OF LEAD, or Sugar of Lead ($\text{Pb}_2\text{C}_2\text{H}_3\text{O}_2$), much resembles loaf sugar in taste and appearance, and is prepared by dissolving litharge in vinegar; it is very poisonous.

CARBONATE OF LEAD—(PbCO_3)—occurs native along with galena. It may be prepared by adding carbonate of soda to a solution of the acetate or nitrate of lead. A highly basic carbonate of lead (sometimes considered as a mixture of lead carbonate and hydrate) is much used as a pigment and for

other purposes under the name of **white lead**. It is prepared on a large scale by exposing plates of lead to the combined action of carbonic acid, moisture, acetic acid and air, aided by a gentle heat. Water, unless charged with carbonic acid gas, does not dissolve it.

SULPHIDE OF LEAD—Galena, (PbS)—is the chief ore of the metal; it may be formed artificially by fusing lead with sulphur, or by precipitating any soluble lead salt by sulphuretted hydrogen.

CHLORIDE OF LEAD (PbCl_2) is best prepared by adding hydrochloric acid to a solution of nitrate of lead: a white precipitate, soluble in about 33 parts of boiling water, is obtained.

IODIDE OF LEAD (PbI_2) is obtained as beautiful yellow spangles, by adding iodide of potassium to a solution of nitrate of lead, dissolving the precipitate in hot water, and allowing it to cool.

TESTS FOR LEAD.

1. Sulphuretted hydrogen, and sulphide of ammonium, give a **black precip.** of the sulphide.

2. Caustic potash gives a **white precip.** of the hydrated oxide, soluble in excess.

3. Iodide of potassium gives a **yellow precip.** of iodide of lead.

4. Bichromate of potassium gives a **yellow precip.** of chromate of lead.

5. Hydrochloric acid gives a **white precip.** of chloride of lead, which is little soluble in cold water, but readily soluble in boiling water.

Blow-pipe test. Salts of lead on charcoal give a soft bluish white bead of lead, surrounded with a yellow incrustation of the oxide.

(3) SILVER (Argentum).

Symbol Ag. *Combining weight* 108. *Sp. gr.* 10.53.

Silver is found native, and in combination with sulphur, antimony, copper, iron, and gold. Galena, the sulphide of lead, usually contains small quantities of it. The mines of Peru, Mexico, Nevada, Kongsberg in Norway, Schneeberg and Freiberg in Saxony, are the chief sources of silver.

The process of extracting the metal from its ores is complicated, and differs according to the kind of ore, and the custom of the country where it is found. Silver ores are not smelted like those of iron, but another method, called **amalgamation** (*i.e.*, combination with metallic mercury), is employed to extract the metal. In Germany, the usual method is to roast, at low red heat, the powdered ore along with salt. By this means the sulphide of silver is converted into the chloride. The mass is then put into casks made to revolve, and scraps of iron and water are added. During the agitation a chemical reaction takes place between the iron and the silver chloride, by which metallic silver and chloride of iron are formed. Mercury is then added, and the mixture again agitated, and a liquid amalgam of mercury and silver (and gold, if any be present) is formed. This amalgam is strained, and then heated, when the mercury distils off, and impure silver remains.

Pure silver may be obtained by dissolving a silver coin in nitric acid, when nitrate of silver is formed. The solution will be greenish-blue from the presence of copper. A solution of common salt is added to the nitrate, and chloride of silver is precipitated. The chloride is washed and melted in a crucible along with carbonate of sodium,—the result is pure silver and common salt.

Properties. Pure silver is white, and of a brilliant lustre. It is highly malleable and ductile, and is the best conductor of heat and electricity known. In hardness, it lies between copper and gold. It fuses at about 1000° C. It does not

oxidize in air, but if melted, it absorbs 22 times its bulk of oxygen, which, however, it expels on cooling. Cold nitric acid and boiling oil of vitriol dissolve it readily. Silver blackens slowly in the air, on account of the presence of sulphuretted hydrogen, which readily combines with the metal, and forms the sulphide. The brilliancy of silver makes it useful for the covering of reflectors in lighthouses. Silver stains glass an orange colour.

ALLOYS OF SILVER. The most common of these is an alloy of silver and copper, for coinage and jewellers' work. Copper is added to give silver sufficient hardness for extensive use. In English standard silver, the amount of copper is 7.5 per cent.

COMPOUNDS OF SILVER.

OXIDES OF SILVER. There are three: a **suboxide** (Ag_4O), which is a black powder, easily decomposed, and soluble in ammonia; a **protoxide** (Ag_2O), which is the base of the silver salts; and a **peroxide** (Ag_2O_2).

The **Protoxide of Silver**, or **Argentite Oxide** (Ag_2O), is obtained as a brown powder by adding caustic potash to a solution of any silver salt. Heat and light decompose it.

The two most important compounds are the **nitrate** and **chloride** of silver.

Nitrate of Silver, or **Argentite Nitrate** (AgNO_3), may be obtained as crystals, by dissolving silver in nitric acid, and evaporating the solution. These are soluble in an equal weight of cold water. When fused, the nitrate is cast into moulds, and then forms the sticks of **lunar caustic**. This salt, in contact with organic matter, is blackened by sunlight. It is used in the preparation of marking ink for linen.

Chloride of Silver (AgCl) is found native as **horn silver**. It is obtained as a dense, curdy, white mass, by adding a solution of any chloride to a soluble salt of silver. When heated it melts, and on cooling crystallizes into what is called **horn**

silver. Sunlight changes it to a violet colour, but if organic matter be present, the chloride is instantly blackened. This property makes it useful in photography. This salt is insoluble in water. Hyposulphite of sodium dissolves it, and on account of this it is used for "fixing" photographic pictures—that is, giving them a permanent character by removing the unaltered salt of silver.

The other compounds, as the **Sulphide** (Ag_2S), **Bromide** (AgBr), **Iodide** (AgI), **Fluoride** (AgF), and **Sulphate** (Ag_2SO_4), do not require any special mention.

The soluble salts of silver are colourless, and act as irritant poisons.

TESTS FOR SILVER.

1. Sulphuretted hydrogen, or sulphide of ammonium, gives a **black** precip. of the sulphide, insoluble in alkalis or dilute acids.

2. Caustic potash gives a **brown** hydrated oxide, insoluble in excess.

3. Hydrochloric acid gives a **white curdy** precip., insoluble in nitric acid, but **soluble** in ammonia.

Blow-pipe test. Heated with carbonate of soda on charcoal, a **white** bead of silver is obtained.

(4) MERCURY (Hydrargyrum).

Symbol Hg. *Combining weight* 200. *Sp. gr. as a liquid*, 13.6.

Mercury or quicksilver is found native in small quantities, but it usually occurs as **cinnabar** (the sulphide of mercury), an ore which is found in Spain, California, Transylvania, China, and Japan.

The metal may be extracted by simply roasting the ore, by which the sulphur is burnt off, and the mercury distilled; or by heating the ore along with slaked lime.

Properties. Mercury is the only metal liquid at the ordinary temperature. It has the lustre of silver. It freezes at 39.4°C .

and is then malleable ; it boils at 350°C ., and forms a colourless vapour, with a sp. gr. of 6.976. It does not tarnish in air, if pure ; but heated to a temperature slightly above its boiling point it slowly combines with oxygen, and forms the red oxide (HgO). At ordinary temperatures it combines directly with chlorine, bromine, iodine, and sulphur. Nitric acid dissolves it with the evolution of nitric oxide ; hydrochloric acid does not affect it ; sulphuric acid, if heated with it, forms mercuric sulphate, with the evolution of sulphurous anhydride (SO_2).

Mercury unites with various other metals either at ordinary temperatures or when slightly heated ; the combinations thus formed are called **amalgams**. On this account mercury is largely used in the extraction of gold and silver from their ores, the mercury being subsequently separated from the other metal by distillation. An amalgam of tin and mercury is used for the reflecting surface of mirrors.

Mercury and its compounds are employed in medicine, while the metal is used in the construction of barometers, thermometers, and other philosophical instruments.

COMPOUNDS OF MERCURY.

Mercury forms two classes of compounds, in one of which it acts to some extent as a monad, while in the other compounds it plays the part of a dyad element. The former compounds are frequently distinguished as **mercurous**, the latter as **mercuric** compounds.

OXIDES OF MERCURY. There are two, a black suboxide (Hg_2O), and a red oxide (HgO), both of which form salts.

The **Suboxide of Mercury**, or **Mercurous Oxide** (Hg_2O), a dark powder, is obtained from **calomel** (the chloride of mercury) by digesting it with caustic potash. Heat and the sun's rays convert it into the red oxide and the metal. It is insoluble in water, but soluble in acids, forming mercurous salts.

The **Red Oxide of Mercury**, or **Mercuric Oxide** (HgO), may be obtained as a red crystalline body by gently heating the

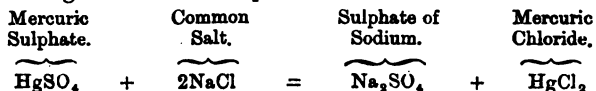
nitrate, or by heating the metal to about its boiling point for some time, or by adding caustic potash to a solution of **corrosive sublimate** (dichloride of mercury, HgCl_2). In the last method the oxide has a yellowish colour. This compound is readily decomposed by heat, yielding oxygen gas and metallic mercury. This oxide is highly poisonous, slightly soluble in water, but dissolves in acids, forming mercuric salts.

There are two sulphides corresponding to the two oxides: a **subsulphide** or **mercurous sulphide** (Hg_2S), which is of little importance, and a **sulphide** or **mercuric sulphide** (HgS). This latter is **cinnabar**, the most abundant ore of mercury; it is also known as the pigment **vermilion**, which is prepared by heating together sulphur and mercury. Mercuric sulphide is precipitated as a black body by sulphuretted hydrogen, but on drying it assumes a red colour.

There are two **chlorides**, which are very important compounds; they are commonly known as **calomel** and **corrosive sublimate**.

The **Chloride of Mercury**, **Mercurous Chloride**, or **Calomel** (Hg_2Cl_2), is obtained by precipitating a solution of mercurous nitrate by the aid of common salt, or by heating a mixture of three parts of finely divided mercury with four parts of corrosive sublimate. Calomel is a heavy, white, tasteless powder, insoluble in water. The alkalies and lime water decompose it. Sulphuric acid has no action upon it; boiling nitric acid, and hydrochloric acid, after long boiling, decompose it. It is largely used for medicinal purposes.

The **Bichloride of Mercury**, **Mercuric Chloride**, or **Corrosive Sublimate** (HgCl_2) is prepared on the large scale by heating a mixture of equal parts of common salt and of mercuric sulphate. It is also obtained by heating metallic mercury in chlorine gas. The changes in the former process are thus shown:—



Corrosive sublimate is highly poisonous, fuses at 265° C., and is soluble in water, alcohol, and ether. In cases of poisoning, white of egg is the usual antidote, as corrosive sublimate forms with this an insoluble compound.

Mercury forms with iodine **three iodides**, and by heating the metal along with oil of vitriol, **sulphate of mercury**, or **mercuric sulphate** (HgSO_4), is obtained.

Nitrates of Mercury, corresponding to the two oxides, are prepared by the action of dilute nitric acid upon the metal, and by dissolving mercuric oxide in nitric acid respectively.

TESTS FOR MERCURY.

In a solution of a mercurous salt :—

1. Sulphuretted hydrogen gives a **black precip.** of subsulphide of mercury.
2. Hydrochloric acid gives a **white precip.** of calomel, insoluble in excess, but soluble in hot nitric acid.
3. Caustic potash gives a **black precip.** of the suboxide.
4. Iodide of potassium gives a **yellowish-green precip.** of mercurous iodide, soluble in excess.

In a solution of a mercuric salt :—

1. Sulphuretted hydrogen, or ammonium sulphide, gives a **dirty white precip.** of the sulphide, changing through red to black, insoluble in excess or in acids.
2. Hydrochloric acid gives **no precip.**
3. Caustic potash gives a **bright yellow precip.** of the oxide.
4. Iodide of potassium gives a **brilliant scarlet precip.** of the iodide; soluble in excess.

Salts of mercury in solution give a silvery metallic coating to clean strips of copper, which disappears on heating. Heated with carbonate of sodium in a tube, they deposit metallic globules.

(5) BISMUTH.

Symbol Bi. *Combining weight* 210. *Sp. gr.* 9.79.

Bismuth is not an abundant metal; it is found native, and also combined with sulphur, and in connection with the ores of nickel and cobalt. It is easily separated by heating the ores in iron tubes.

It is a hard, brittle, metal, of a pinkish-white colour, and fuses at 264° C. It does not oxidise in dry air at the ordinary temperature, but if heated strongly it burns with a blue flame, forming an oxide (Bi_2O_3). When thrown in a powdered state into chlorine gas it takes fire, forming the trichloride (BiCl_3). Nitric acid dissolves it rapidly, while hydrochloric and sulphuric acid have little action on it. It is chiefly used in making the alloy known as fusible metal, which consists of 2 parts of bismuth, 1 of lead, and 1 of tin, and melts at about 94° C., or below the temperature of boiling water. Its compounds are used for pigments and in medicine.

OXIDES OF BISMUTH are two in number,—a sesquioxide, or bismuth oxide (Bi_2O_3), which is a yellow, insoluble powder, obtained by heating the nitrate; and a pentoxide, or bismuthic oxide (Bi_2O_5), a brown powder, obtained by dissolving the hydrated sesquioxide in potash, and transmitting chlorine gas.

The most important soluble salt of bismuth is the **nitrate**, or bismuth nitrate ($\text{Bi}(\text{NO}_3)_3 + 5\text{H}_2\text{O}$), which is prepared by dissolving the metal in nitric acid.

TESTS FOR BISMUTH.

1. Sulphuretted hydrogen gives a **black** precip. of the sulphide (Bi_2S_3); insoluble in dilute acids, and in alkalis.
2. Caustic potash gives a **white** precip., insoluble in excess, which becomes **yellow** on boiling.
3. Water gives a **white** precip. with soluble bismuth salts, and this reaction is very characteristic. The precip. is soluble in an acid.

With the **blow-pipe**. Bismuth compounds, heated on charcoal with carbonate of sodium, give a metallic bead with an incrustation of the oxide, which is reddish-yellow while hot, but becomes pale-yellow on cooling.

(6) CADMIUM.

Symbol Cd. *Combining weight* 112. *Sp. gr.* 8.6.

Cadmium is usually found in the zinc ores as sulphide of cadmium. Being more volatile than zinc, cadmium distils over first in the preparation of the former metal. It resembles tin in appearance; it is malleable and ductile, and fuses at about 315° C. It does not oxidise in air, but when made red-hot it takes fire, depositing a brown oxide (CdO). Cold nitric acid, and hot sulphuric or hydrochloric acid, dissolves it readily.

The cadmium salts are colourless, and resemble those of zinc.

TESTS FOR CADMIUM.

Sulphuretted hydrogen gives a **yellow** precip. of the sulphide, insoluble in alkaline solutions, or in ammonium sulphide.

With the **blow-pipe**. Cadmium compounds, heated on charcoal with carbonate of soda, give a brown incrustation of oxide of cadmium.

DIV. III. *Metals which give no precipitate with Sulphuretted Hydrogen in alkaline solutions:—Gold, Platinum, Tin, Antimony, Arsenic, &c.*

(1) GOLD (Aurum).

Symbol Au. *Combining weight* 196.2. *Sp. gr.* 19.3.

Gold is a metal widely diffused, and always found in the metallic state in the primitive rocks, and in the beds of certain rivers. The most celebrated mines are those of California, Australia, Mexico, Chili, Brazil, and Peru.

Gold which occurs in sand is extracted by the simple process of washing in a "cradle," and forms the "gold-dust" of com-

merce. Rocks containing gold are crushed to powder, and the metal extracted by the process of amalgamation with mercury.

Properties. Gold has a beautiful yellow colour, and, when pure, is almost as soft as lead. It is eminently malleable and ductile, and fuses at about 1200°C . It is a good conductor of heat and electricity. It does not tarnish in air, and it resists the action of every single acid, excepting selenic. A mixture of one part of nitric acid, and four parts of hydrochloric acid, called *aqua regia*, dissolves it; so also does any mixture which liberates chlorine.

Pure gold may be obtained by dissolving the ordinary commercial metal in *aqua regia*, and adding ferrous sulphate, when the gold is precipitated as a brown powder, which becomes bright by friction.

Gold is extensively used for ornaments, &c., but on account of its softness, it requires to be mixed or alloyed with other metals. Jewellers' gold is usually mixed with copper and silver. British **standard gold** is an alloy of eleven parts of gold to one of copper.

OXIDES OF GOLD. Of these there are two,—a **suboxide** (Au_2O), and a **trioxide** (Au_2O_3). Neither of these forms salts with acids; but the latter, when hydrated, has acid properties, and is sometimes called **auric acid**. It forms with bases, salts called **aurates**. The oxides of gold are very unstable compounds, undergoing decomposition by the action of heat alone.

The most important compounds of gold are the two chlorides, the **protochloride** (AuCl), and the **trichloride** (AuCl_3).

The **Protochloride**, or **Aurous Chloride** (AuCl) is obtained as a pale-yellow insoluble powder, by heating the trichloride to about 175°C ., and thus expelling two-thirds of its chlorine.

Trichloride of Gold, or **Auric Chloride** (AuCl_3) is the most important gold compound, and is obtained by dissolving the metal in *aqua-regia*, and evaporating to dryness. The trichloride is a red crystalline body, very deliquescent, and soluble

in water, alcohol, and ether. It unites with alkaline chlorides, and forms double salts, which are mostly of a yellow colour when in crystals, and red when deprived of water.

The remaining gold compounds are of no importance.

TESTS FOR GOLD.

1. A solution of **protosulphate of iron** gives a brown precip. of metallic gold.

2. **Chloride of tin** gives a purple precip., called **purple of cassius**.

(2) PLATINUM.

Symbol Pt. *Combining weight* 197.1. *Sp. gr.* 21.5.

Platinum is a rare metal, always found native, and mixed with the metals palladium, rhodium, osmium, ruthenium, and iridium. Sometimes it is found mixed with gold, silver, copper, iron, and lead. Our chief supply comes from the Ural Mountains, and smaller quantities are obtained from Mexico, Brazil, and Australia.

The process of obtaining platinum is a complicated one. The crude metal is dissolved in aqua-regia, and precipitated in combination with several of the other metals by means of sal-ammoniac (ammonium chloride). This precipitate of **ammonic-platinic chloride** is heated, and produces a spongy mass called **spongy platinum**, which, when heated and pressed, forms a metallic mass.

A new method of extracting platinum has been lately introduced. The ore is melted in a furnace made of lime, heated with the oxy-hydrogen blow-pipe; most of the impurities are expelled, and an alloy of platinum, iridium, and rhodium is obtained, which is more useful than the pure metal for most purposes, being harder and less fusible than pure platinum.

Properties. Platinum has a white colour, and is capable of a high lustre; hence it was called by the Spaniards of South America *platina* (little silver). It does not tarnish in air, and

can only be melted by the oxy-hydrogen blow-pipe, or the heat of the electric arc. It is malleable and ductile. It resists the action of all single acids; but aqua-regia dissolves it. If heated to redness with the caustic alkalies, it is oxidised, and combines with the alkaline base.

On account of its infusibility, and its power of resisting chemical agents, platinum is largely used for the construction of chemical apparatus.

Platinum and oxygen do not unite together directly, but two oxides can be formed indirectly, viz., the monoxide (PtO), and the dioxide (PtO_2).

The most important compounds of platinum are the chlorides, viz., the protochloride, or platinous chloride (PtCl_2); and the perchloride, or platonic chloride (PtCl_4).

The former is obtained by evaporating a solution of platinum in aqua-regia, and exposing the residue to a heat of 204°C . It is a dark green powder, insoluble in water.

The perchloride is the most important salt, and is obtained by evaporating a solution of the metal in aqua-regia. It dissolves in water, forming an orange-coloured solution in water, and it forms double salts with the chlorides of the alkaline metals, (*e.g.* ammonic platonic chloride, $2(\text{NH}_4)\text{Cl}, \text{PtCl}_4$).

TESTS FOR PLATINUM.

1. Sulphuretted hydrogen and sulphide of ammonium both give a black precip. of the sulphide, soluble in excess of sulphide of ammonium, insoluble in nitric or hydrochloric acid.

2. Caustic potash gives a yellow precip., soluble in excess.

(3) TIN (Stannum).

Symbol Sn. *Combining weight* 118. *Sp. gr.* 7.3.

Tin is always found in the state of ore, running in veins in the primary rocks, and is usually mixed with the sulphides of copper and iron. Its chief ore is *tinstone*, which is the dioxide of tin, and is found in Cornwall, Saxony, Banca, and Malacca.

The metal is extracted by first crushing and washing the ore, to separate it as far as possible from earthy impurities, —then by roasting, to expel sulphur, &c.,—and then by reduction and refining. English tin usually contains small quantities of arsenic, copper, iron, and lead.

Properties. Tin is, next to silver, the whitest of the metals. It is soft, ductile, and malleable, but has little tenacity. On account of its malleability, it may be hammered or rolled out into very thin sheets, known as tin-foil. It is a good conductor of heat and electricity; it fuses at 230° C. It tarnishes very slowly when exposed to air at the ordinary temperature; if exposed to great heat, it takes fire and burns, forming the dioxide. Hot hydrochloric acid dissolves tin, with the liberation of hydrogen, and the formation of stannous chloride; dilute nitric acid converts it into an insoluble oxide, known as meta-stannic acid, and nitrous fumes are evolved.

Tin is largely used to protect the more oxidisable metals (iron, &c.) on account of its property of remaining practically unaltered in both dry and moist air.

ALLOYS OF TIN are numerous. **Tin-plate** consists of thin plates of iron coated with tin: **Britannia metal** consists of a variable amount of tin combined with equal parts of brass, antimony, and bismuth: **Pewter** consists of four parts of tin and one of lead: **Queen's metal** is an alloy similar to Britannia metal. With copper, tin forms several alloys, as **speculum metal**, **bell metal**, **gun metal**, and **Bronze**. With mercury, tin forms an amalgam for the silvering of looking-glasses. Plumbers' solder consists of 30 to 60 per cent. of tin, the remainder being lead. **Type metal** is an alloy of equal parts of antimony and tin, with two of lead.

COMPOUNDS OF TIN.

OXIDES OF TIN. There are two oxides,—a protoxide, or stannous oxide (SnO), and a dioxide, or stannic oxide (SnO_2).

Stannous Oxide (SnO) is an olive-brown powder, obtained by

heating the hydrated protoxide ($\text{Sn}_2\text{H}_2\text{O}_3$) in carbonic acid. The hydrated protoxide is obtained by adding a solution of a stannous salt to a strong solution of an alkaline carbonate.

The Dioxide, or Stannic Oxide (SnO_2), occurs native as **tin-stone**, but it may be prepared artificially in two forms, having very different properties. When tin is acted upon by nitric acid, a white powder is produced, which is insoluble in acids, and is called **meta-stannic acid**. If an alkaline solution be added to a solution of bichloride of tin, a white precipitate of stannic oxide is formed, soluble in acids; this is distinguished from the former variety by the name of **stannic acid**.

There are two chlorides, the **dichloride of tin**, or **stannous chloride** (SnCl_2), and the **tetrachloride of tin**, or **stannic chloride** (SnCl_4). The former, in a hydrated state, is obtained by dissolving tin in hydrochloric acid; it is used by the dyer and calico printer as a "mordant," that is, a substance used for the purpose of fixing the colouring matter in the material.

Stannic chloride is prepared by passing chlorine gas over metallic tin; it is a thin, colourless liquid, emitting dense white fumes when exposed to the air. When mixed with water, it forms a solid crystalline hydrate. It is largely used in dyeing for fixing colours.

TESTS FOR TIN.

(1). *In an acid solution of a protosalt of tin* :—

1. Sulphuretted hydrogen, or sulphide of ammonium, gives a **dark brown** precip. of sulphide of tin (SnS), which is soluble in solutions of potash or soda, and in yellow ammonium sulphide.
2. Caustic potash gives a **white** precip. of the hydrated stannous oxide, soluble in excess.
3. Chloride of gold gives a **purple** precip. (the purple of cassius).

(2). *In an acid solution of a persalt* :—

1. Sulphuretted hydrogen, or sulphide of ammonium, gives a **dull yellow** precip. of stannic sulphide (SnS_2), soluble in excess, and in alkalis.

2. Caustic potash gives a white precip., soluble in excess.

Before the blow-pipe, salts of tin form with carbonate of soda, in the inner flame, white malleable beads of metallic tin, and emit white fumes of stannic dioxide.

(4) ANTIMONY (Stibium.)

Symbol Sb. *Combining weight* 122. *Sp gr.* 6.71.

Antimony is an abundant mineral, chiefly found in combination with sulphur, as stibnite, or grey antimony ore (Sb_2S_3). The metal is easily extracted from this ore by heating it with iron, when the sulphur and iron combine, and the antimony is set free; or by mixing the ore with coal, and heating in a reverberatory furnace.

Properties. Antimony is a bluish-white metal, very brittle, and highly crystalline. It fuses at 430°C . At ordinary temperatures it does not oxidise in air; but if heated, it burns readily with a white flame, giving off white fumes of antimonious oxide (Sb_2O_3). Powdered antimony takes fire when thrown into chlorine gas. Aqua-regia readily dissolves the metal, so also does hot strong hydrochloric acid. Nitric acid attacks the metal, and converts it into a white insoluble oxide.

Antimony is much used in alloys, as mentioned under the heads of Lead and Tin.

COMPOUNDS OF ANTIMONY.

OXIDES OF ANTIMONY. There are only two oxides of any importance, viz.,—trioxide or antimonious oxide (Sb_2O_3); and pentoxide or antimonie oxide (Sb_2O_5).

The Antimonious Oxide, or Trioxide of Antimony (Sb_2O_3), may be prepared by adding a solution of carbonate of soda to a solution of trichloride of antimony, when the oxide is precipitated as a yellow powder. When boiled with bitartrate of potassium (cream of tartar), it dissolves, and yields, on evaporation, crystals of tartar emetic. This oxide is the basis of the salts of antimony used in medicine. It dissolves in hydro-

chloric and tartaric acids. Nitric acid converts it into a higher oxide, and sulphuric acid forms with it an insoluble sulphate.

The **Antimonic Oxide**, also called **Antimonic Acid** (Sb_2O_3), is obtained by the action of nitric acid upon antimony. It is a pale yellow powder, insoluble in water, and loses part of its oxygen on the application of heat. Salts, called **antimoniates**, are known in which this oxide acts the part of an acid.

Antimony has a strong affinity for chlorine, and forms with it two chlorides, viz., **antimonious chloride**, or **trichloride of antimony** (SbCl_3), and **antimonic chloride**, or **pentachloride of antimony** (SbCl_5). Of these the first is the most important. It may be prepared by dissolving sulphide of antimony in hydrochloric acid, and distilling the liquid. The trichloride passes over, and forms, on cooling, white crystals. It is used for bronzing gun barrels.

Antimony forms, with sulphur, two compounds corresponding to the oxides, viz., **antimony trisulphide** (Sb_2S_3), and **antimony pentasulphide** (Sb_2S_5). The former occurs native, as **grey antimony ore**, possessing a bluish black colour, and metallic lustre; and it is the only ore from which the metal is obtained.

Antimony combines with hydrogen, and forms a gaseous compound called **antimoniuretted hydrogen** (H_3Sb), which is colourless, and burns with a bluish flame, forming white fumes of antimonious oxide. This gas is produced when sulphuric acid is added to zinc, in presence of a soluble compound of antimony. Metallic antimony is deposited by the flame on a piece of cold porcelain, brought in contact with it, forming a bright metallic mirror.

TESTS FOR ANTIMONY.

1. Sulphuretted hydrogen gives, in an acid solution, an orange precip. of antimonious sulphide, which is soluble in sulphide of ammonium, and in alkalies.

2. If the above precipitate be washed and gently dried, and

a few drops of hydrochloric acid added, trichloride of antimony is formed, which, when thrown into water, gives a white precipitate.

Blow-pipe test Compounds of antimony, mixed with carbonate of soda, give in the inner flame globules of the metal, and dense white fumes of antimony trioxide. The metal may be dissolved in hot hydrochloric acid, and tested as above.

(5) ARSENIC.

Symbol As. Combining weight 75. Sp. gr. 5.7 to 5.9.

Arsenic is sometimes classed amongst the non-metallic elements. It resembles phosphorus in its chemical properties. It is sometimes found native, but it is a very common constituent of the ores of iron, cobalt, nickel, copper, or tin.

The metal is extracted from its ores by roasting them in a furnace, and condensing in flues the trioxide of arsenic (As_2O_3) thus produced. This oxide is further purified and sold as **white arsenic**. The powdered oxide, mixed with charcoal, is heated in a retort,—the metal passes off in vapour, and is condensed in a cool part of the apparatus.

Properties. Arsenic possesses a steel-grey colour and brilliant lustre, but it oxidises in moist air, and becomes tarnished. It is crystalline and very brittle. When heated it passes, without fusing, into a vapour, which is colourless, and smells strongly of garlic. It burns in air with a bluish flame, from which pass off fumes of the trioxide. Thrown into chlorine gas, it takes fire, and is changed into the trichloride of arsenic. Nitric acid acts on the metal, and changes the metal into arsenic pentoxide (As_2O_5), but hydrochloric acid has little action upon it.

Gun shot consists of lead and a little arsenic. The trioxide is largely used in preparing green and yellow paints, and also in the manufacture of flint glass.

COMPOUNDS OF ARSENIC.

OXIDES OF ARSENIC. There are two oxides—the trioxide, or **arsenious acid** (As_2O_3), and the pentoxide, or **arsenic acid** (As_2O_5).

Arsenious acid—white arsenic (As_2O_3)—is produced by burning the metal in air, and by roasting the ores. It exists in the form of crystals, and also as a glassy substance. It is slightly soluble in water, and dissolves in the alkalis, forming with them soluble compounds called **arsenites**. This oxide forms with copper the pigment called **Scheele's green**, which is a strong poison.

Arsenic acid (As_2O_5) is obtained by the action of nitric acid upon arsenious acid. It is a white powder, soluble in water; if the solution be evaporated, crystals of true arsenic acid (H_3AsO_4) are formed. This oxide is decomposed by heat into arsenious acid and oxygen.

Arsenic acid is a powerful acid, and forms with metals salts corresponding to the phosphates, called **arseniates**. The arseniates of the alkalis are soluble in water; those of the other metals are insoluble, but are dissolved by the acids. This oxide is a less powerful poison than arsenious acid.

Arsenic and sulphur form three compounds, viz., **realgar** (As_2S_2); the **trisulphide** or **orpiment** (As_2S_3); and the **pentasulphide** (As_2S_5). The first two compounds occur native; the trisulphide is deposited as a bright yellow powder, when sulphuretted hydrogen is sent through a solution of arsenious acid in hydrochloric acid ($2\text{AsCl}_3 + 3\text{H}_2\text{S} = 6\text{HCl} + \text{As}_2\text{S}_3$).

Arsenic also forms compounds with chlorine, bromine, and iodine. Like antimony, it combines with hydrogen, and forms **arseniuretted hydrogen** (AsH_3), which is a gaseous body of some importance, corresponding to phosphoretted hydrogen. It may be prepared by decomposing an alloy of arsenic and zinc with dilute sulphuric acid. The greatest care is needed in the preparation of this gas, as the inhalation of one bubble

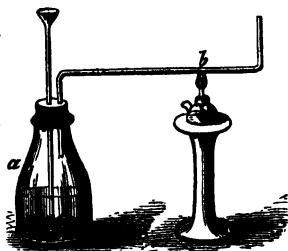
has been known to cause death. It burns with a bluish white flame, producing arsenious acid; if a cold body be held in the flame, metallic arsenic is deposited upon it. Chlorine decomposes this gas with flame; it is also decomposed by sending it through a red-hot tube. At a temperature of 40° below zero, the gas is reduced to a colourless liquid.

TESTS FOR ARSENIC.

1. Sulphuretted hydrogen gives in an acid solution a **bright lemon-yellow precip. of orpiment**, soluble in ammonia, and in ammonium sulphide.

2. Sulphate of copper gives in neutral solutions a **green precip. of arsenite of copper** (Scheele's green).

Marsh's Test. This test depends upon the production of arseniuretted hydrogen when hydrogen is prepared in the presence of soluble compounds of arsenic. Zinc and dilute sulphuric acid are put into a flask (a) having a funnel tube, and a delivery tube bent at right angles. The arsenical liquid is added by means of the funnel, and arseniuretted hydrogen is given off. If a lamp be held at (b), the gas, in passing, is decomposed, and deposits metallic arsenic on a portion of the tube beyond the flame. Or the jet of hydrogen may be burnt at the end of the tube, and a cold piece of porcelain held in the flame, when, if arsenic (or antimony) be present, a bright metallic mirror will be formed on the porcelain. The arsenical mirror is soluble in a solution of sodium hypochlorite; that formed by antimony is insoluble. In employing this test it is necessary that the zinc and sulphuric acid be pure.



In testing with the **blow-pipe** the presence of arsenic is rendered evident by the peculiar garlic-like odour of its oxide.

The remaining metals, on account of their rarity, require but

little notice. Some of their compounds are used in the arts and manufactures.

URANIUM.

Symbol U. Combining weight 240. Sp. gr. 18.4.

This metal is sparingly distributed, and is found combined in the two minerals, pitchblende (uranium tetroxide, U_3O_4), and uranite. It is closely allied in its chemical relations to iron and manganese. It has a steel white colour; does not oxidize in air at ordinary temperatures; burns brilliantly when heated; and dissolves in sulphuric and hydrochloric acids. It forms several oxides and salts. The uranium compounds are useful for glass staining, and photography. The protoxide, or uranous oxide (UO), gives a black colour to glass, whilst the sesquioxide, or uranic oxide (U_2O_3), colours glass a beautiful greenish yellow.

INDIUM.

Symbol In. Combining weight 113.4. Sp. gr. 7.36.

This metal is found in certain ores of zinc, and was discovered by means of spectrum analysis. It is a white malleable metal, slowly soluble in hydrochloric acid; its compounds give to flame a blue colour.

TITANIUM.

Symbol Ti. Combining weight 48.

This is a very rare metal, analogous to tin in its chemical characters. Titanic oxide (TiO_2) is found native under the names of rutile and anatase in Cornwall and other localities. It forms several compounds, but none of them are used in the arts.

MOLYBDENUM.

Symbol Mo. Combining weight 96. Sp. gr. 8.6.

This is a white, brittle metal, the chief ore of which is the disulphide, (molybdenite, MoS_2), found in Bohemia and Sweden. Molybdenite of lead (or wulfenite, $PbMO_4$) occurs in Hungary

and other localities, in the form of beautiful yellow crystallized plates. It forms three oxides; one of these, molybdic trioxide (MoO_3), acts upon bases, and yields salts, called molybdates. The molybdate of ammonia is employed as a delicate test for small quantities of phosphoric acid, in solutions acidulated with nitric acid; a yellow liquid is produced, which, on gently heating, deposits a yellow precipitate. The compounds of molybdenum are not used in the arts.

TUNGSTEN (Wolfram).

Symbol W. Combining weight 184. Sp. gr. 17.4.

This metal is found combined with lime in the mineral scheelite (CaWO_4), and with ferrous oxide in wolfram (FeWO_4). It has been obtained only as a dark-grey powder, which is very hard, and difficult of fusion. A little of the metal mixed with steel gives the latter extraordinary hardness. Tungsten forms several compounds. Tungstate of soda is sometimes used in the laundry in connection with starch, for the purpose of rendering muslin, &c., incombustible.

THALLIUM.

Symbol Tl. Combining weight 204. Sp. gr. 11.9.

This metal was discovered by spectrum analysis in 1861, in the deposit formed in the flue of a furnace in which iron pyrites was burned for the sulphuric acid manufacture. In physical properties it resembles lead. It has a bluish-white lustre on fresh surfaces, but it oxidises rapidly on exposure to the air; it is soft, malleable, and ductile. It melts at 290°C .; if heated higher, it burns with a brilliant green flame. It burns in chlorine, and combines directly with bromine, iodine, sulphur, and phosphorus. Nitric and sulphuric acids dissolve it readily, with evolution of hydrogen. With hydrochloric acid it forms an insoluble chloride. It forms several oxides and salts; the latter are poisonous. In combination, it usually acts as a

monad, as shown by its sulphide (Tl_2S), and chloride ($TlCl$), but there is a higher class of compounds in which thallium acts as a triad, e.g. ($TlCl_3$).

The remaining metals, **Niobium** or **Columbium**, **Tantalum**, **Vanadium**, **Palladium**, **Rhodium**, **Ruthenium**, **Osmium**, **Iridium**, require no special mention.

GENERAL QUESTIONS ON THE METALS PROPER.

1. Describe the method of extracting Zinc from its ores. How would you detect its presence in a solution? (C.)
2. What are the chemical differences between Wrought Iron, Cast Iron, and Steel? Explain Bessemer's process for the manufacture of Cast Steel. 1.5 grammes of pure iron wire are ignited in excess of Oxygen Gas: what is the nature, and what the weight of the resulting compound? (C.)
3. What are the properties of the metal Platinum? How is it prepared? Name its most important salt, and state its use in the laboratory. (C.)
4. A piece of green paper is supposed to contain Arsenic. How would you try to ascertain whether it does contain it or not? (O.)
5. The atomic weights of Silver and Chlorine being 108 and 35.5, determine the quantity of metal contained in 14 grains of Chloride of Silver. (O.)
6. What reactions will prove the presence of the following Bases in a solution—(1) Lime, (2) Alumina, (3) Cobalt, (4) Sesquioxide of Iron, (5) Peroxide of Tin? (O.)
7. If 50 grains of ignited Oxide of Copper be exposed to a current of Hydrogen until its reduction is completed, how much Copper will be obtained? The atomic weight of Copper is 31.75. (O.)
8. Describe the metallurgic process by which Lead is obtained from its most common ore. (O.)
9. Ten grains of a mixture of Chloride of Silver and Iodide of Silver contain $1\frac{1}{2}$ grains of Chlorine. How much Iodine and Silver are present? The atomic weights of Chlorine, Iodine, and Silver are 35.5, 127, and 108. (O.)
10. Explain what is meant by an equivalent of a substance. If 100 parts of Oxide of Copper, reduced by Hydrogen, give 22.69 of water, deduce the equivalent of Copper? (C.)
11. Explain how pure Silver may be made from an alloy of that metal with Copper. (C.)

12. Explain how Sulphate of Copper is prepared, and how Copper may be detected in a mixed solution. (C.)

13. State and explain the effect of a solution of Hydrochloric Acid upon each of the following: (1) Protoxide of Lead, (2) Red Oxide of Iron, (3) Black Oxide of Manganese, (4) Sulphide of Antimony, (5) Silver, (6) Nitric Acid. (C.)

14. How are (1) Cobalt, (2) Arsenic, and (3) Manganese recognised by the Blow-pipe? (O.)

15. What reactions are characteristic of the presence of the following metals in a solution: (1) Copper, (2) Bismuth, (3) Lead? (O.)

16. The atomic weight of Iron being 28, determine the quantity of metal contained in 100 grains of Sesquioxide of Iron. (O.)

17. By what means may Metallic Silver be obtained from its Chloride? (O.)

18. From what ore is Copper most commonly procured? Describe the process of its reduction. Express by symbols its compounds with Oxygen. (C.)

19. A person at a distance from a town needs some Nitrate of Silver for analytical purposes: he has pure Nitric Acid in his laboratory, how can he improvise the required re-agent? (C.)

20. What are the compounds of Iron and Oxygen? What are the chemical characters of the Salts of Iron? (C.)

21. Define—(1) Metal, (2) an Alloy. What is the composition of (i) Brass, (ii) Pewter, (iii) German Silver? (C.)

22. In what form does Chromium occur in nature? How would you convert Chromic Acid into Sesquioxide of Chromium? What are the chemical characters of these substances? (C.)

23. Describe Marsh's test for Arsenic, and state the precautions necessary in using it. (C.)

24. 100 grains of a substance are found to contain Iron 63.6, Sulphur 36.4. Find its chemical formula, $Fe=28$, $S=16$. (C.)

25. Describe any of the methods adopted in the manufacture of Alum. What is its formula? (C.)

26. What are the principal ores of Iron, and in what districts are they found? Describe the method by which Cast Iron is converted into Wrought Iron. (C.)

27. What changes do the following substances undergo when heated—Oxide of Silver, Nitrate of Copper, Nitrate of Ammonium, Sulphate of Iron, Oxalate of Calcium, Chloride of Platinum and Potassium? (O.)

28. How are the following metals affected when heated to redness in the air: Iron, Copper, Silver, Bismuth, Tin? (O.)

29. What are the sources, and what is the mode of preparation of the metal Aluminium? (O.)

30. What takes place when Red Lead (Minium) is boiled with Nitric Acid? *

31. What is the composition of Arseniетted Hydrogen? and how is it prepared? *

32. Give the name and composition of the mineral in which Mercury occurs in nature, and describe the process by which the metal is obtained? *

33. The Copper is to be precipitated from 100 grammes of Crystallized Sulphate of Copper ($\text{CuSO}_4 + 5\text{H}_2\text{O}$) by means of Metallic Zinc. How much Zinc is required? *

34. You are required to convert 1 lb. of Silver into Argentic Nitrate (Nitrate of Silver)—how will you do it, and what will be the weight of Nitrate produced? *

35. 1,000 gra. of Copper Oxide are placed in a glass tube, which is then heated, and a current of pure hydrogen is passed through the tube; what will be the chemical action, and what the weights of the substances severally produced? ($\text{H} : \text{O} : \text{Cu} = 1 : 16 : 63.5$). (C.)

36. Two powders are given you: one is a mechanical mixture, the other a chemical combination, of Sulphur and Iron. In what respects would the two powders differ, and by what chemical reaction could you distinguish one from the other? (L.)

37. Soda-water has not unfrequently been found to contain small quantities of Lead in solution. How would you ascertain, with certainty, whether this substance was present, and how would you endeavour to determine its amount? (L.)

38. A powdered mineral when mixed with carbonate of soda, and heated on charcoal before the blowpipe, gives a bright, metallic, malleable bead. This, when boiled with nitric acid, yields an insoluble white powder. What metal do these reactions indicate, and what confirmatory tests would you use? *

39. By what characters are metals distinguished from non-metals?

40. State in words, and give an equation to illustrate, the action of sulphuric acid upon zinc, copper, and common salt respectively.

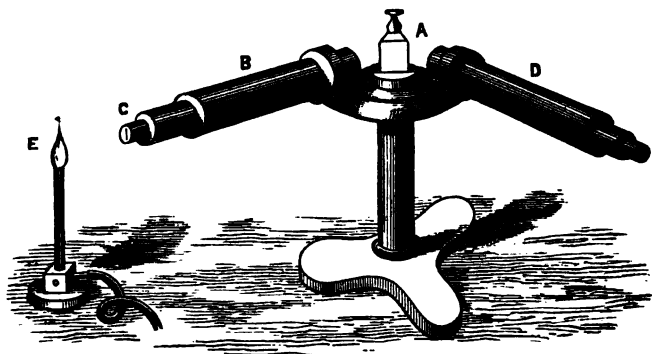
LIGHT AND SPECTRUM ANALYSIS.

Sir Isaac Newton discovered that a ray of sunlight consisted of several coloured waves of light. He allowed a sunbeam to pass into a darkened room through a hole in the shutter of the window. A triangular piece of glass, called a **prism**, was placed near the hole. The ray of light, in passing through the prism, was broken up into its constituent parts, which appeared upon a screen in a long band of various colours, according to the order of their refrangibility. This coloured band is called a **spectrum**, and the order in which the colours appear, beginning at the least refrangible, is as follows :—**red, orange, yellow, green, blue, indigo, violet**. Each source of pure white light (*e.g.* the electric light, the lime light, &c.) gives a continuous spectrum similar to that of the sun.

Dr. Wollaston, in 1802, made a further discovery with regard to the solar spectrum. By allowing the light to pass through a magnifying lens, before falling upon the prism, he found certain gaps, or dark lines, crossing the spectrum in different parts at right angles to its length. These lines were first carefully mapped by an optician of Munich, named **Fraunhofer**, and they are still called after him, **Fraunhofer's lines**. For a long time these lines were a mystery to scientific men, but they at length received explanation from observing the spectra of various coloured flames. It has long been known that certain chemical substances, when heated strongly in the blow-pipe flame, or flame of the Bunsen burner, will impart to that flame a peculiar colour characteristic of those substances. Thus, sodium compounds give the flame an intense yellow colour; while potassium compounds impart a violet colour. If a ray of light from these flames be examined in the same manner as a sunbeam mentioned above—*viz.*, by letting the thin beam of light, which has passed through a narrow slit, fall on a prism—there will be seen on the screen, not a variously coloured band like the solar spectrum, but a spectrum of a few bright

lines only. Sodium salts, for instance, will give a single broad line of yellow light; potassium salts, two bright lines of red and violet respectively; hydrogen gives three bright lines, and so on. As every substance capable of being heated to the point of luminosity gives its own peculiar spectrum, it is very easy to determine, by an examination of the spectrum, the presence or absence of the substance in question. In this way the presence of the most minute quantity of a substance can be detected. This delicate method of analysis is called **spectrum analysis**. Now, if the solar spectrum be compared with that of any of the substances mentioned above, it will be found that the bright bands peculiar to each of the latter exactly correspond to the gaps or dark lines in the former. For instance, the yellow band characteristic of the sodium salts, exactly corresponds to a dark line in the solar spectrum; from which fact we infer that incandescent sodium emits a ray wanting in sunlight. It is evident, therefore, that there exists a real connection between the metals which emit bands corresponding to the Fraunhofer lines, and the materials of which the sun is composed. This subject of scientific enquiry has been carefully investigated by the famous chemists Bunsen and Kirchhoff of Heidelberg, and the latter in 1859 established the following fact, that, "When light from a luminous solid or liquid passes through a gas, the gas absorbs those particular rays of light of which its own spectrum consists." Thus, then, the mystery of the Fraunhofer lines was explained. In the sun's atmosphere there exist the vapours of various metals, and as the ray of sunlight emitted from the solar body passes through these, they absorb the particular rays of light which they themselves emit when incandescent, and occasion the gaps or dark lines on the solar spectrum. This explanation of the Fraunhofer lines, therefore, determines what substances are present in the sun's atmosphere; and by an examination of the spectra of the stars, it can be determined what substances are present in those heavenly bodies.

An instrument called a **spectroscope** has been devised for the examination of light obtained from various sources.



It consists of a stand, upon which the prism (A) is fixed in a vertical position. On one side is fixed a telescope (B), the eyepiece of which is replaced by a tube containing two brass plates, which, by means of a screw, can be brought nearly close together so as to form a very narrow slit (C). On the other side is an ordinary telescope (D). The light from a substance held in the flame of the Bunsen burner (E), passing through the slit, is refracted into parallel rays by means of the telescope (B), and fall upon the prism (A). The rays are then received by the telescope (D), and magnified before reaching the eye of the observer. A black cloth is thrown over the prism and telescopes to exclude extraneous light.

A striking proof of the value of this instrument lies in the fact that several elementary bodies have been discovered by its use. Bunsen, in 1859, by the spectrum analysis of the residue of the water of a German mineral spring, determined the existence of the alkaline metals Cesium and Rubidium. The spectrum of the first metal was characterised by two blue lines, and that of the second by two red lines. Shortly after the

metal Thallium was discovered by Mr. Crookes, who observed in the spectrum of a residue from a German chemical manufactory, a splendid green line, which did not belong to the spectrum of any known substance. Later still, the element Indium was discovered by its spectrum of a fine dark blue line.

For further information on this very important and interesting subject, the student is referred to Roscoe's **Lectures on Spectrum Analysis**.

HINTS ON QUALITATIVE ANALYSIS.

ANALYSIS is the name given to that branch of Chemistry which has for its object the discovery of the chemical **composition** of bodies. It is sub-divided into two branches, viz, **Qualitative** and **Quantitative** Analysis; the former discovers the **nature** and properties of the component parts of a compound, and the latter determines the **quantity** of each present.

The ordinary way of analysing a simple compound,—i.e., a body containing only **one** base and **one** acid, or **one** metal and **one** non-metallic element,—is to mix some **reagent** or **test** with a solution of the substance to be analysed. The **reagent** is so called because, when placed in contact with another substance, it produces a particular appearance or change, termed the **reaction**, by which the presence of a particular substance is shown.

The following are the chief operations involved in Qualitative Analysis:—

1. **SOLUTION.** This term denotes the perfect union of a solid with a fluid. The liquid which dissolves the body is called the **solvent**. Solutions are of two kinds—**simple** or **mechanical**, and **chemical**. A solution is simple when no chemical change has taken place, as a solution of salt and water; it is chemical when an entirely new substance is formed, consequent upon chemical combination, as chalk dissolved in hydrochloric acid.

When the solvent has received as much of the solid as it can contain, the solution is said to be **saturated**.

2. **PRECIPITATION.** This is a term given to the formation of an insoluble compound of the body held in solution, by the action of a **reagent**, or **precipitant**. The insoluble substance thus formed is called a **precipitate**. Precipitates are classified according to their appearance, as **curdy**, **crystalline**, **gelatinous**, &c. Precipitates are called **cloudy**

or turbid, when they are so fine and so small as only to affect the transparency of the fluid. The colour of a precipitate with a certain reagent is usually a sign of the substance previously in solution.

3. **FILTRATION.** This is the separation of the solid matter suspended in a liquid, by passing the latter through suitable filtering paper. The liquid which passes through the filter is called the **filtrate**.
4. **DECANTATION.** This is the separation of heavy solid particles in a fluid, by simply inclining the vessel and pouring the liquid off after the solid has sunk to the bottom.
5. **EVAPORATION.** This is the process of obtaining solid matter from solutions, by evaporating the latter by the aid of heat. If the heat be gently applied, the solid matter frequently takes a crystalline form, and the process is then called **crystallization**.
6. **DISTILLATION** is the term given to the process of separating a liquid from a solid, or one liquid from another having a different boiling point, by the application of heat. The volatile liquid passes in a gaseous form into a receiver, and is there condensed into the liquid state.
7. **IGNITION** does for solid bodies what evaporation does for liquids: it separates the **volatile** and **non-volatile** parts of a solid. It is performed by heating the substances in vessels called crucibles, or on porcelain, or platinum foil.
8. **SUBLIMATION** is the term given to the distillation of a solid. The volatile part of the solid is driven off by heat, and is recondensed in a receiver by cold. The substance is then called a **sublimate**.
9. **FUSION** is the conversion of a solid into the liquid form by heat. When an insoluble solid is fused along with some other substance, in order to form a soluble compound, the process is called **fluxing**, and the body employed for this purpose is termed a **flux**.
10. **DEFLAGRATION** is a term applied to all decompositions

attended with a noise or detonation. Nitrates and chlorates deflagrate when heat is applied to them in the "dry state."

11. **THE BLOW-PIPE.** The use of this instrument is a most important operation. A description of the instrument and its uses have been given at page 50.

LIST OF APPARATUS REQUIRED FOR ANALYSIS.

Bunsen gas burner.

Retort stand.

Filter stand.

Test tube stand.

Mouth blow-pipe.

Sand bath.

Wire gauze.

Crucible tongs.

Spatula.

Platinum wire.

Platinum foil.

Test tube brushes.

Test tubes, 6 in. \times $\frac{3}{4}$ in.

German glass flasks, 2 oz., 4 oz., and 8 oz

Glass beakers.

Glass funnels, 2 in. and 3 in. diam.

Thistle funnels.

German glass retorts, stoppered.

Berlin porcelain crucibles, $1\frac{1}{2}$ in. diam.

Berlin porcelain evaporating basins, $2\frac{3}{4}$ in. and $3\frac{1}{2}$ in.

Test glasses.

Glass tubing, soft, $\frac{3}{16}$ in. to $\frac{1}{4}$ in. diam.

Thin glass rod, for stirring.

Filter paper, cut to suit funnels.

Corks, assorted.

Cork borers.

Caoutchouc tubing, $\frac{3}{8}$ in. and $\frac{1}{2}$ in. bore.

Triangular file to cut glass.

Glass mortar and pestle.

Wash bottle.

Woulff's bottles for preparing carbonic acid, &c.

1 spirit lamp, 4 oz.

Methylated spirit.

Pieces of charcoal for blow-pipe experiments.

The following reagents will also be required :—

Sulphuric acid.

Nitric acid.

Hydrochloric acid.

Tartaric acid.

Lime water.

Ammonia (solution).

Caustic potash.

Sodium carbonate.

Barium chloride.

Potassium ferrocyanide.

Silver nitrate.

Ammonium oxalate.

Ammonium sulphide.

Ammonium carbonate.

Ammonium chloride.

Ferric chloride.

Ferrous sulphide (for sulphuretted hydrogen).

Platinic chloride.

Calcium sulphate.

Sodium phosphate.

Potassium iodide.

Potassium nitrate (crystals).

Acetate of lead.

Zinc.

Copper.

Starch.

Blue and red litmus test papers.

Borax

Cobalt nitrate (solution)

DIRECTIONS FOR THE ANALYSIS OF SIMPLE COMPOUNDS.

1. The substance to be examined should be reduced to a state of fine powder. Examination in the dry state by means of the blow-pipe, &c., is described in Table I.
2. In making a solution of the substance, the usual solvents tried are water, hydrochloric acid, nitric acid, and aqua-regia, in the order named. First, put a little of the powdered body in a test-tube with a little **distilled water**. If the substance fails to dissolve in the cold, hold the tube over the spirit-lamp, and boil. If it is still undissolved, add **gradually** one of the above acids, beginning with hydrochloric.
3. Examine first for the **base**, and be careful to use the reagents **gradually**, and also consider for what purpose you add each particular reagent. Table II. shows the classification of the metallic bases, and the tests by which the various groups are distinguished. Tables III., IV., V., VI., VII., furnish tests for each base.
4. After finding the base, examine for the **acid** by means of Table VIII. Some of the acids are found in examining for the base.
5. In using the reagents, if a precipitate is not formed at once, stir with a glass rod, or gently heat the solution.
6. Use the substance to be analysed with economy. Always keep a portion for confirmatory tests.
7. Be careful to note every change during the process of examination, and immediately make a record of it.
8. The test-tubes, &c., must be perfectly clean.
9. Needless exposure to injurious gases, as chlorine, sulphuretted hydrogen, &c., should be avoided.

[f.] Ammonia " Ammoniacal Salts.

[g.] Cyanogen, known by the } Cyanides.
smell of bitter almonds,

[h.] The substance turns black, } An Acetate or Tar-
evolving vapours having } t stg, or other organic
the smell of burnt paper, } salt.

6. The substance fuses A salt of the Alkalies, or Baric, Strontic, or Calcic
to a clear transparent Chloride.
liquid.

III.

Heat the substance on Charcoal with Carbonate of Soda
in the inner flame.

Metallic reduction is observed.

May be

Bismuth	Brittle metal, yellow incrustation.
Antimony	{ Ditto, volatilizes in white fumes with abundant incrustation.
Silver	{ Malleable, very white, no incrustation on charcoal, soluble in Nitric Acid.
Tin.....	{ Malleable, slight white incrustation, con- verted by Nitric Acid into white Meta- stannic Acid.
Lead	{ Very malleable: marks paper, yellow incrustation.
Copper } Gold	{ Characteristic colour of the metal.
Iron } Cobalt } Nickel }	{ Dark infusible metallic powders; magnetic.

Zinc Oxide, green. Observe the colour
of the outer-flame.
Aluminium } blue.
Silica }
Phosphates }
Magnesium, pink.
Strontium, crimson
Calcium } brownish
red.
Sodium, yellow.
Potassium, violet.

IV.

Heat the substance on Platinum Wire with a bead of Borax.

Coloured beads are given in the

Outer-flame.	Inner-flame.
Copper, greenish-blue.	Copper { (If much of the sub- stance has been added, red and opaque.) colourless.
Cobalt, blue.	Cobalt, blue.
Nickel, brownish-red.	Nickel, grey and turbid.
Iron { orange-red while hot, pale yellow when cold.	Iron, bottle-green when cold.
Manganese, amethyst.	Manganese { colourless or cloudy.
Chromium, green.	Chromium, green.
Bismuth } light yellow. Silver }	Bismuth } colourless. Silver }

TABLE II.
CLASSIFICATION OF METALLIC BASES

Group I.	Group II.	Group III.	Group IV.	Group V.
Precipitate produced by Hydrochloric Acid.	Precipitate produced by Sulphuretted Hydrogen, in acid solution.	Precipitate produced by Sulphide of Ammonium after adding Chloride of Ammonium and Ammonia.	Precipitate produced by Carbonate of Ammonium in presence of Chloride of Ammonium.	Bases remaining after the separation of the former groups.
—	—	—	—	—
METALS.	Division I.	Division II.		
LEAD—white.	Sulphides insoluble in Sulphide of Ammonium.	Sulphides soluble in Sulphide of Ammonium.	BARUM—white.	MAGNESIUM.
SILVER—white.	—	—	STRONTIUM—white.	POTASSIUM.
MERCURY—white	MERCURY—black.	TIN—{ brown or yellow.	CALCIUM—white.	SODIUM.
	LEAD—black.	ANTIMONY—{ orange red.		AMMONIUM.
	BISMUTH—black.	ARSENIC—yellow.		
	COPPER—black.			
	CADMIUM—yellow.			
	PLATINUM—{ brownish black.			
	GOLD—{ brownish black.			
				N.B.—Phosphate of Sodium in the presence of Ammonia separates Magnesium.

TABLE III.

BASES—GROUP I.

To the solution add HCl in excess. If there be a precipitate, it is either silver chloride (AgCl), mercurous chloride (Hg_2Cl_2), or lead chloride (PbCl_2).

To distinguish between these, take a small part of the wet precipitate, and pour on it an excess of ammonia (NH_4HO).

Re-dissolves, Ag .	Turns black, Hg.	Is unaltered, Pb.
<p>Confirm.</p> <p>Add Nitric Acid (HNO_3); the precipitate reappears.</p>	<p>Confirm.</p> <p>A drop of the original solution placed on a bright slip of copper (Cu), will, in a few minutes, produce a spot which, when gently rubbed, appears shining like silver. If the copper is gently heated, the mercury (Hg) volatilizes, and the silvering disappears.</p> <p>Also heat original substance with sodium carbonate in glass tube; globules of mercury will condense on sides.</p>	<p>Confirm.</p> <p>The pp. is soluble in hot water, crystalline needles of PbCl_2 separate from solution as it cools.</p> <p>The dried precipitate heated on charcoal before the blow-pipe, gives malleable metallic globules.</p> <p>Sulphuric Acid (H_2SO_4) produces a white precipitate of sulphate of lead (PbSO_4).</p>

TABLE IV.
BASES—GROUP II.

To the solution, treated with HCl, add Sulphuretted Hydrogen, until it smells strongly after shaking.

If only a white precipitate (sulphur) be formed, proceed as if there were none. If there be a coloured precipitate, it may be one of the following: PbS , HgS , Bi_2S_3 , CuS , CdS , SnS , SnS_2 , Sb_2S_3 , As_2S_3 , Au_2S_3 , PtS_2 .

To determine which of these is present, observe the colour.

Yellow.		Orange.		Brown or Black.	
<p>May be CdS, As_2S_3, or SnS_2.</p> <p>To determine which mix a portion of the liquid (containing the precipitate), with excess of NH_4HO, add 4 or 5 drops of $(NH_4)HS$, and warm.</p>		<p>Sb_2S_3.</p> <p>—</p> <p>Confirm by heating original substance on charcoal in blowpipe flame.</p>		<p>PbS, HgS, Bi_2S_3, CuS, SnS.</p> <p>Test the original solution (a small portion at a time) with the following, (except in the last case, when the dark precipitate is used):—</p>	
Undissolved.	<p>As_2S_3, SnS_2.</p> <p>Reprecipitate by adding dilute HCl, and digest pp. with ammonium carbonate.</p>	<p>H_2SO_4.</p> <p>A white precipitate.</p> <p>Fb.</p>	<p>KHO.</p> <p>A yellow precipitate.</p> <p>Hg.</p> <p>Confirm as in Table III.</p>	<p>NH_4HO in excess.</p> <p>A white precipitate; dissolve the precipitate in a drop or two of HCl; and then add plenty of water. The solution becomes milky.</p> <p>Bl.</p>	<p>NH_4HO in excess.</p> <p>If on adding NH_4HO the precipitate at first formed is redissolved, leaving a blue solution.</p> <p>Cu.</p> <p>Confirm.</p> <p>Test original solution with potassium ferrocyanide — a brownish red precipitate.</p>
	<p>Disolved.</p> <p>As.</p> <p>Confirm by heating original substance with sodium carbonate in closed tube. Metallic mirror and smell of As.</p>				
<p>Cd.</p> <p>Confirm.</p>	<p>NH_4HO in excess, and then 5 or 6 drops of $(NH_4)HS$, to the dark pp. and warm.</p> <p>The precipitate will dissolve if it is</p> <p>Sn.</p> <p>Confirm.</p> <p>Add to this solution some diluted HCl, a light yellow precip.</p>				

TABLE V.
BASES. GROUP III.

To the original solution add Chloride of Ammonium and Ammonia, and, whether the latter reagent produce a precipitate or not, a little Sulphide of Ammonium, and apply heat. A precipitate shows the presence of one of the following: Fe, Co, Ni, Mn, Zn, Cr, Al, or Silicic Acid.

Observe the Colour.

Black.		Flesh-coloured.	Bluish Green.	White.	
<i>May be the Sulphides of Fe, Co, or Ni.</i>		Mn.	Cr.	<i>May be Al, Zn, or Silicic Acid.</i>	
To distinguish these, add to the original solution Potash or Soda hydrate.		Confirm by heating with Sodium Carbonate, and Nitre, on platinum foil.	Confirm by borax bead; and by fusing on platinum foil, with Carbonate of Soda and Nitre.	To distinguish these, add to the original solution a few drops of Sodium Hydrate. A precipitate which dissolves in excess indicates—	
Dirty greenish white precip., changing to red-dish brown on exposure to air.	Light green precip. NI.	Sky-blue, turning to light red on boiling.		Zn or Al.	No precipitate.
	Confirm with Ammonia.	Co.		Pass SH ₂ .	SiO ₂ .
Fe.		Confirm with Blow-pipe.	A bright green mass, soluble in water.	A white precipitate.	No precipitate.
Confirm with Ferrocyanide of Potassium.				Zn.	Al.
				Confirm with Blow-pipe and Nitrate of Cobalt.	Confirm with Blow-pipe and Nitrate of Cobalt.

TABLE VI.
BASES. GROUP IV.

To the original solution, add Chloride of Ammonium, Carbonate of Ammonium and Ammonia, and heat gently.

A white precipitate indicates Carbonate of Barium (BaCO_3), Strontium (SrCO_3), or Calcium (CaCO_3).

To distinguish these, add to the original solution, Sulphate of Calcium.

No precipitate, Calcium. Confirm with Oxalate of Ammonium.	A precipitate after some time	
	Strontium. Confirm by flame colouration.	Barium. Confirm by flame.

TABLE VII.
BASES. GROUP V.

(A). To the original solution, mixed with Chloride of Ammonium, Carbonate of Ammonium, and Ammonia, add **Phosphate of Sodium**, and stir with a glass rod.

A white precipitate indicates **Magnesium** (Ammonium Magnesium Phosphate)

(B). Add to the original solution—

(1.) Tartaric Acid.		(2.) Caustic Potash, and heat gently. Ammoniacal odour. NH_4^+	Confirm by flame.
White pp. K or NH_4^+ .	No pp. Na .		
		K	Confirm by flame.

TABLE VIII. TESTING FOR THE PRINCIPAL ACIDS.

A considerable amount of information as to the acid constituent of a compound substance may be obtained during the Preliminary Examination for the base. Thus Nitrates and Chlorates deflagrate when heated on charcoal. Charring of the substance indicates a Tartrate or Citrate. Chromic Acid and Silicic Acid will have been detected during the examination for the base. It is also well to remember that the following classes of salts are as a rule soluble in water:—Nitrates, Chlorides (except AgCl and Hg₂Cl₂;—PbCl₂ is soluble in boiling water only); Sulphates (except BaSO₄, SrSO₄, PbSO₄;—CaSO₄ is only slightly soluble); whereas among compounds insoluble in water we have Oxides and Sulphides (except Oxides and Sulphides of Alkaline, and Alkaline Earth Metals); Carbonates and Phosphates (except Carbonates, and Phosphates of Alkalies).

I. Heat some of substance with dilute HCl. Effervescence occurs with Carbonates (CO₃), Sulphides (SH₂), Sulphites (SO₃)—SO₂ detected by odour; SH₂ detected by odour and blackening of paper moistened with solution of acetate of lead; CO₂ detected by lime water.

II. Heat substance with concentrated H₂SO₄.—Blackening indicates presence of organic matter (Tartrates, Citrates, &c.) Chlorides give pungent white fumes. MnO₂ added evolves Cl₂. Bromides give red fumes of Bromine. Colour starch paper yellow.

(b.) Iodides give violet vapours of Iodine. " " blue.

(d.) Nitrates give off acid vapours of HNO₃; and if Copper added, red fumes of N₂O₄ are evolved. Confirm also by gently adding to substance acid vapours of FeSO₄; a brown ring at junction of liquids indicates Nitric Acid.

(e.) Fluorides emit pungent acid fumes which etch glass.

(f.) Acetates give off pungent vapours of Acetic Acid.

(g.) Borates on adding Alcohol and igniting, produce a green flame.

III. Boil the substance for some time with Sodium Carbonate, so that the acid may be transferred to the Sodium. Filter and test successive portions of the filtrate as follows:—

(1.) Acidify a portion with HCl.

(a.) Add BaCl₂ solution—white pp. indicates Sulphuric Acid.

(b.) Add AmCl and Am₂CO₃—a gelatinous pp. indicates Silicic Acid.

(c.) Add Fe₂Cl₆ solution—deep blue pp. (Prussian blue) indicates Hydroferrocyanic Acid.

(2.) Acidify a portion with HNO₃.

(a.) Add AgNO₃—a white curdy pp., readily soluble in Ammonia, indicates Hydrochloric Acid.

(b.) A curdy-yellow pp., soluble in Ammonia with difficulty, indicates Hydrobromic Acid.

(c.) " " almost insoluble in Ammonia, indicates Hydriodic Acid.

(3.) Acidify with Acetic Acid.

(a.) Add Fe₂Cl₆ solution—a yellowish gelatinous pp. of Phosphate of Iron indicates Phosphoric Acid. Add to alkaline solution Ammonia and Ammonium Chloride, and then Sulphate of Magnesium,—a white crystalline pp. indicates Phosphoric Acid.

APPENDIX TO TABLES OF ANALYSIS.

As Table II necessitates the almost constant use of Hydrogen Sulphide, the following course shows how several metallic bases may be discovered without the use of this reagent.

To a solution of the substance, add a few drops of Sulphide of Ammonium.

A precipitate. Observe the colour.	
White indicates Zinc or Aluminium. To distinguish, add to the original solution Caustic Potash in excess, and then Chloride of Am- monium.	Flesh colour indicates Manganese. Black indicates one of the remaining Metals proper. To distinguish, proceed as in Table II.
<div> <div>No precipitate indicates an Alkali, or an Alkaline Earth.</div> <div> <div>No precipitate indicates one of the Alkalies. To distinguish, Vide Table VII.</div> <div> <div>A white precipitate indicates one of the Alkaline Earths. To distinguish, Vide Table VI.</div> <div> <div>A white precipitate indicates Aluminium.</div> <div> <div>No precipitate indicates Zinc.</div> </div> </div> </div> </div></div>	

To the original solution, add Phosphate of Sodium and Ammonia.

MEASURES OF CAPACITY.					
	In Cubic Inches.	In Cubic Feet = 1,728 Cubic Inches.	In Pints = 34.65923 Cubic Inches.	In Gallons = 8 Pints = 277.47384 Cubic Inches.	In Bushels, = 8 Gallons = 2218.19076 Cubic Inches.
Millilitre, or cubic centimetre	0.061027	0.0000353	0.001761	0.00022010	0.000027512
Centilitre, or 10 cubic centimetres	0.610271	0.0003532	0.017608	0.00220097	0.000275121
Decilitre, or 100 cubic centimetres	6.102705	0.0035317	0.176077	0.02200967	0.002751208
Litre, or cubic decimetre	61.027052	0.0353166	1.760773	0.22009668	0.027512085
Decalitre, or centistère	610.270515	0.3531658	17.607734	2.20096677	0.275120848
Hectolitre, or decistère	6102.705152	3.5316581	176.077341	22.00966767	2.751208482
Kilolitre, or stère, or cubic metre	61027.051519	35.3165807	1760.775314	220.09667675	27.512084854
Myriolitre, or decastère	610270.515194	353.1658074	17607.734140	2200.96676750	275.120848597
1 Cubic Inch = 16.3869769 Cubic Centimetres.					
1 Gallon = 4.546457989 Litres.					
MEASURES OF WEIGHT.					
	In English Grains.	In Troy Ounces = 480 Grains.	In Avoirdupois Lbs. = 7,000 Grains.	In Grains. = 113 Lbs. = 784,000 Grains.	Tons = 20 Cwt. = 14,980,000 Grains.
Milligramme	0.015432	0.000032	0.0000022	0.00000002	0.000000001
Centigramme	0.154323	0.000322	0.0000220	0.00000020	0.000000010
Decigramme	1.543235	0.003215	0.0002205	0.00000197	0.000000098
Gramme	15.432349	0.032151	0.0022046	0.00019688	0.000000984
Decagramme	154.323488	0.321507	0.0220462	0.00196884	0.000009842
Hectogramme	1543.234880	3.215073	0.2204621	0.00196841	0.000098421
Kilogramme	15432.348800	32.150727	2.2046213	0.01968412	0.000984206
Myriogramme	154323.488000	321.507267	22.0462126	0.19684118	0.009842059
1 Grain = 0.064798960 Gramme.					
1 Troy oz. = 31.103496 Gramme.					
1 lb. Avo. = 0.45359235 Kilogr.					
1 Cwt. = 50.80237089 Kilogr.					

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